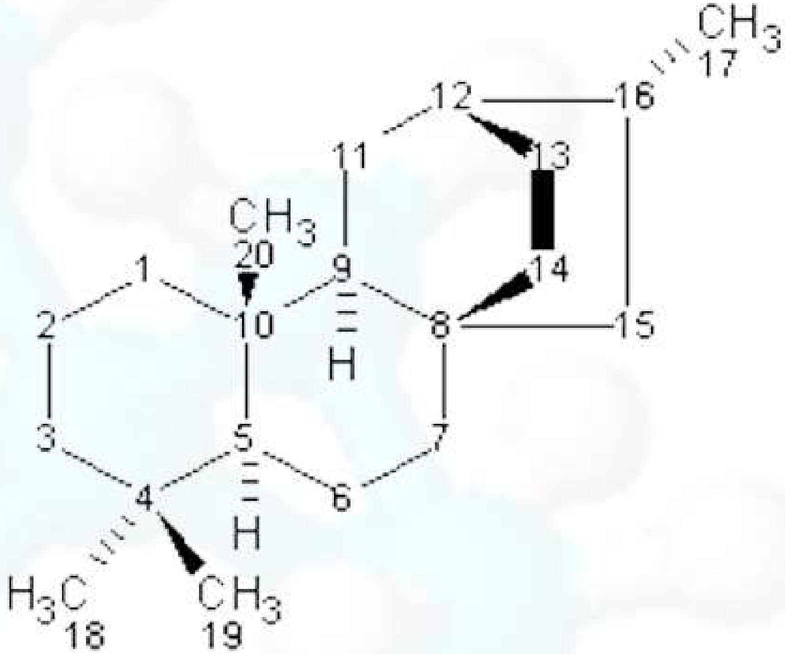


Welcome to
Chem. 110
General Chemistry

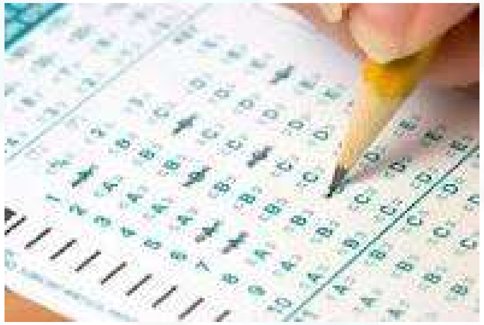
General Information

- Grade:

30% First midterm exam

30% Second midterm exam

40% Final exam



Important notice

If your absence percentage was more than 25% you will be banned from final exam and fail the course.

General Information

- Grade system

A^+ = 100 – 95%

A = 94 – 90%

B^+ = 89 – 85%

B = 84 – 80%

PASS

C^+ = 79 – 75%

C = 74 – 70%

D^+ = 69 – 65%

D = 64 – 60%

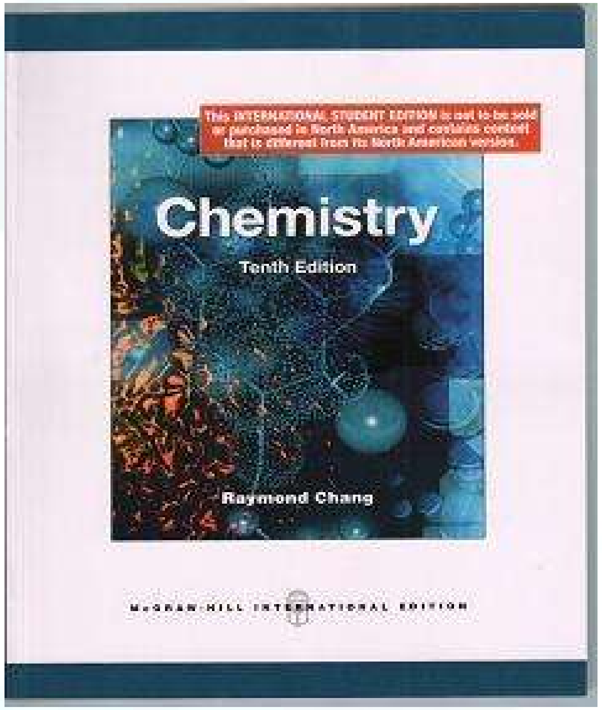
F = <60%

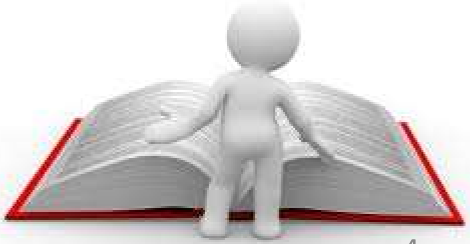
FAIL

Reading Books

- Chemistry by Raymond Chang,
  Tenth Edition.
- Chemistry by Steven S. Zumdahl, Sixth Edition.
- Chemistry by Mortimer, Sixth Edition.
- General Chemistry System by Marwani and Albar.
- Any General Chemistry text book.
- Required chapters:

1 – 2 – 3 – 4 – 5 – 7 – 8 – 9 – 14 – 15 – 16 – 24 – 25





Curriculum

- *Chapter one: Chemical Foundations*
- Units of Measurement
- The Fundamental SI Units
- The Prefixes Used in SI System
- *Chapter two: Atoms, Molecules and Ions*
- The Atomic Theory
- The Structure of the Atom
- Atomic Number, Mass Number and Isotopes
- The Periodic Table
- Molecules and Ions
- Chemical Formulas
- Naming Compounds
- *Chapter Three: Mass Relationships in Chemical Reaction*
- Stoichiometric Calculations: Amount of Reactants and Products
- Avogadro's Number and Molar Mass of an element
- Chemical Equations and Reactions, Balancing Chemical Equations
- Percent Composition of Compounds
- Determining the Formula of a Compound: Empirical, Molecular and Structural Formulas
- Calculation Involving Limiting Reagents
- Molecular Mass
- Atomic Mass
- Reaction Yield

Curriculum

- *Chapter Four: Reaction in Aqueous Solutions*
- Concentration of Solutions: The Molarity, Mole Fraction and Dilution

- *Chapter Five: Gases*
- Substance That Exist as Gasses
- Pressure of a Gas
- The Gas Laws of Boyle, Charles and Avogadro
- The Ideal Gas Equation
- Dalton's Law of Partial Pressure
- *Chapter Seven: Quantum Theory and the Electronic Structure of Atoms*
- From Classical Physics to Quantum Theory
- Bohr's Theory of the Hydrogen Atom
- The Dual Nature of the Electron
- Quantum Numbers   Electromagnetic Radiation
- Atomic Orbitals
- Electron Configuration
- The Building-up Principle
- *Chapter Eight: Periodic Relationships Among the Elements*
- Development of the Periodic Table
- Periodic Classification of the Elements

Curriculum

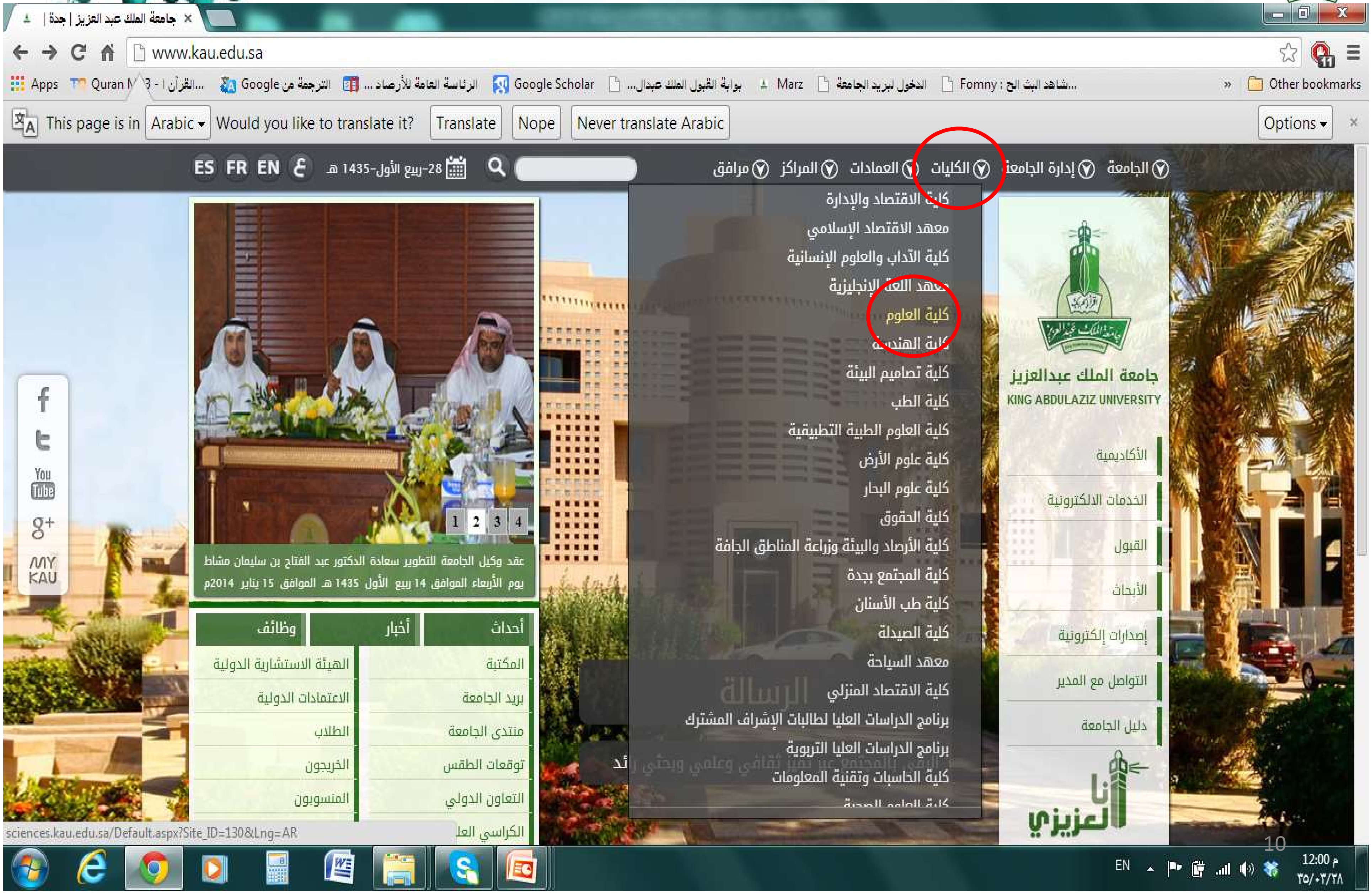
- Periodic Variation in Physical Properties
- Ionization Energy
- Electron Affinity
- *Chapter Nine: Bonding: General Concepts*
- Lewis Dot Symbols
- The Ionic Bond
- Types of Chemical Bonds: Ionic and Covalent Bonds
- Electronegative
- Writing Lewis Structures
- Formal Charge Lewis Structures
- The Concept of Resonance
- Exceptions to the Octet Rule
- *Chapter Fourteen: Chemical Equilibrium*
- The Concept of Equilibrium and the Equilibrium Constant
- Writing Equilibrium Constant Expressions
- The Relationship Between Chemical Kinetics and Chemical Equilibrium
- What Does the Equilibrium Constant Tell Us
- Factors That Affect Chemical Equilibrium

Curriculum

- *Chapter Fifteen: Acids and Bases*
- The Acid-Base Properties of Water
- pH-A Measure of Acidity
- Weak Acids and Acid Ionization Constants
- *Chapter Sixteen: Acid-Base Equilibria and Solubility Equilibria*
- The Common Ion Effect
- Buffer Solutions
- Solubility Equilibria
- Chapter Twenty-four: Organic Chemistry
- Aliphatic Hydrocarbons (Alkane, Cycloalkanes, Alkene, Cycloalkenes and Alkynes, Including the NOMENCLATURE)
- Classes of Organic Compounds
- Geometrical Structures Isomerism FOR ALKENES ONLY *cis-trans* isomers
- Aromatic Hydrocarbons (Including the NOMENCLATURE and *o,m,p*)
- Chemistry of the Functional Groups (Functional Group Only)
- *Chapter Twenty-five: Synthetic and Natural Organic Polymers*
- Proteins
- Nucleic Acids

How To Get A+ In This Course

- Find the lectures in website , print, bring
- In Lectures listen, write , ask
- Practice with old exams and on university websit as much as you can



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عقد وكيل الجامعة للتطوير سعادة الدكتور عبد الفتاح بن سليمان مشاط يوم الأربعاء الموافق 14 ربيع الأول 1435 هـ الموافق 15 يناير 2014م

sciences.kau.edu.sa/Default.aspx?Site_ID=130&Lng=AR

أهلا بكم في موقع كلية العلوم بجامعة الملك عبد العزيز

نبذه عن كلية العلوم

أنشئت كلية العلوم في العام الجامعي 1393\1394هـ الموافق 1973\1977م وكانت تشمل الأقسام العلمية التالية : قسم الرياضيات،قسم الفيزياء،قسم الكيمياء،قسم الجيولوجيا،قسم الأحياء،ثم أنشئ بعد ذلك قسم لعلوم البحار وقسم الأرصاد وشعبة للعلوم الفلكية تابعة لقسم الفيزياء ومع مطلع العام الجامعي 1398\1399هـ انفصل قسم علوم البحار حيث أصبح معهداً مستقلاً قائماً بذاته تابعاً للجامعة وتحت إشرافها ثم عدل اسمه في العام الجامعي 1400\1401هـ إلى كلية علوم البحار كما انفصل في نفس العام قسم الأرصاد وتحول إلى معهد عالي سمي بكلية الأرصاد ودراسات المناطق الجافة الذى عدل اسمه فيما بعد إلى

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كلية العلوم

المواقع التفاعلية الخاصة بمواد السنة التحضيرية بكلية العلوم

عزيزي طالب / طالبة السنة التحضيرية
يسعدنا تلقي الآراء والمقترحات على الايميل aalmozaini@kau.edu.sa , ezz@kau.edu.sa

تعليمات هامة فيما يخص
رياضيات 110 ورياضيات 111

كيمياء 205 كيمياء 110

رياضيات 111 رياضيات 110

احصاء 111 احصاء 110

جامعة الملك عبدالعزيز
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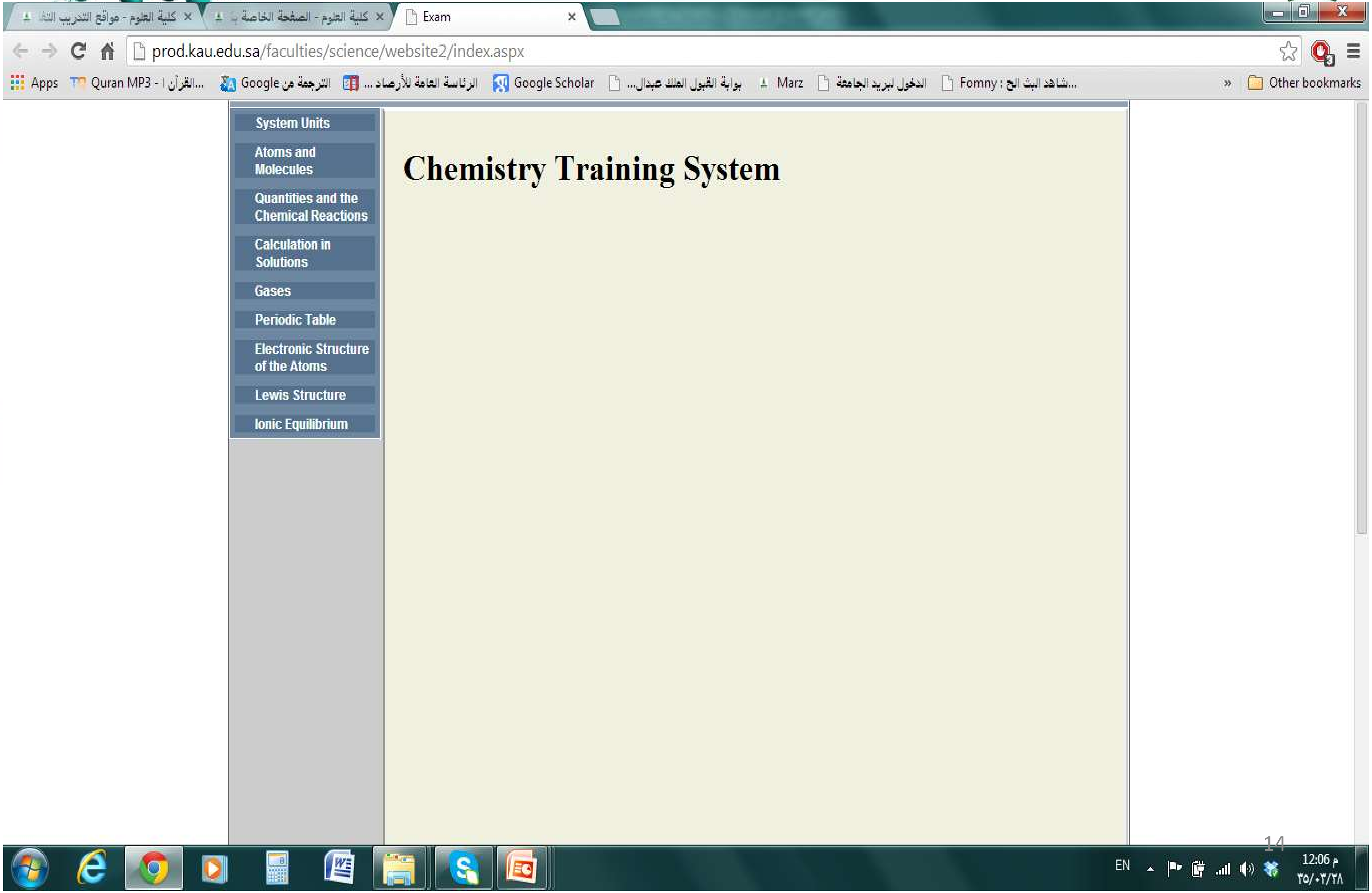
الصفحة الخاصة بكيمياء 110 للسنة التحضيرية

مرحبا بكم في الصفحة الخاصة ب كيمياء 110

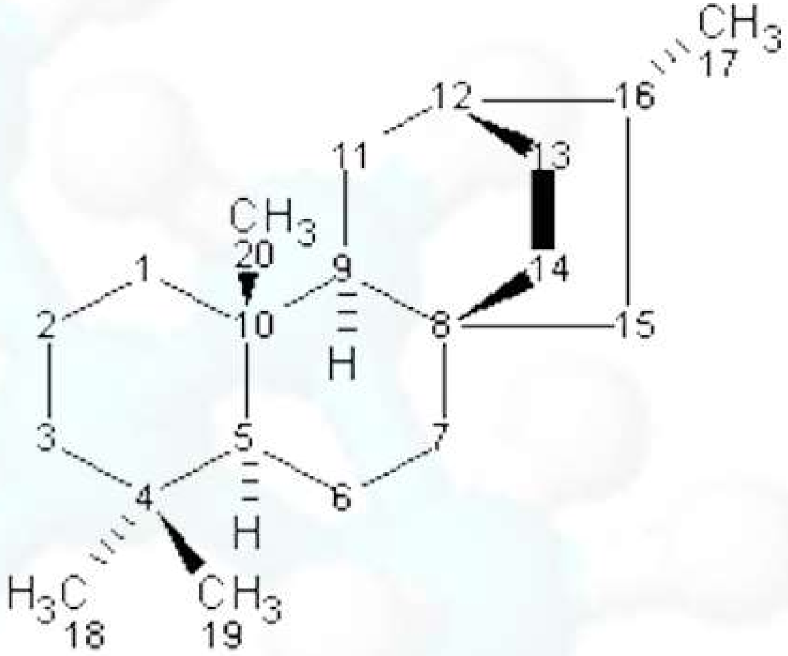
CHEM 110

مع تمنياتنا للجميع بالتوفيق والسداد



Chemistry Training System

Thank you

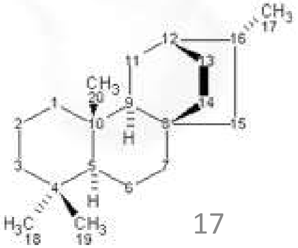


Chapter One

Introduction

Chapter 1 Chemical Foundations <u>(p15-20)</u>

- Chemistry: An overview
- Units of Measurement
- The Fundamental SI Units
- The Prefixes Used in SI System



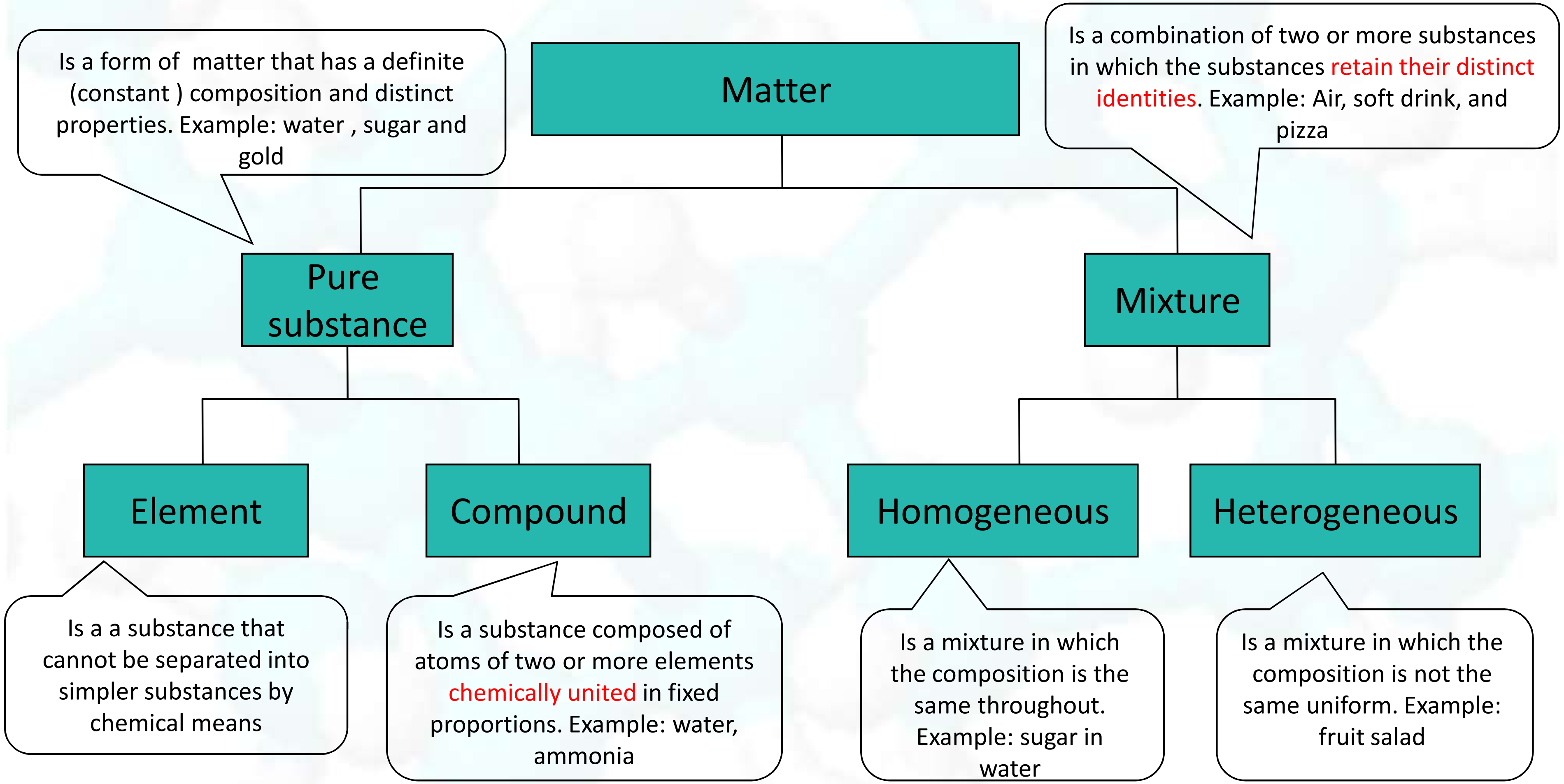
- What is Chemistry?

- Chemistry is the study of Matter's composition, structure, and properties and changes it undergoes.

- Matter is anything that occupies space and has mass

- Thus Matter is everything around us.

Chapter One / Introduction



Is a form of matter that has a definite (constant) composition and distinct properties. Example: water , sugar and gold

Matter

Is a combination of two or more substances in which the substances retain their distinct identities. Example: Air, soft drink, and pizza

Pure substance

Mixture

Element

Compound

Homogeneous

Heterogeneous

Is a a substance that cannot be separated into simpler substances by chemical means

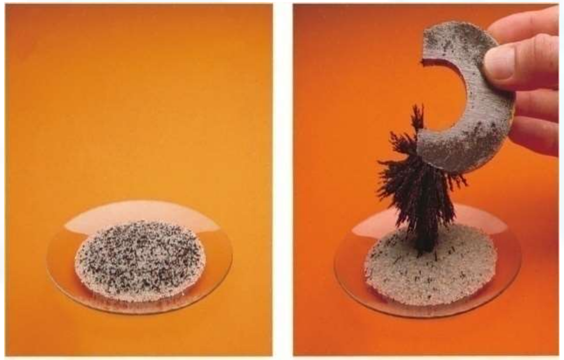
Is a substance composed of atoms of two or more elements chemically united in fixed proportions. Example: water, ammonia

Is a mixture in which the composition is the same throughout. Example: sugar in water

Is a mixture in which the composition is not the same uniform. Example: fruit salad

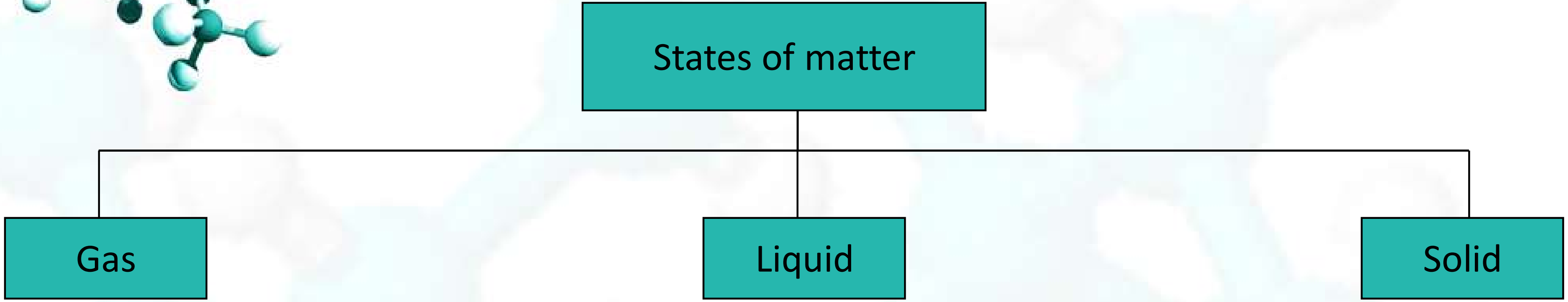
Element

compound



Heterogeneous mixture

Homogeneous mixture

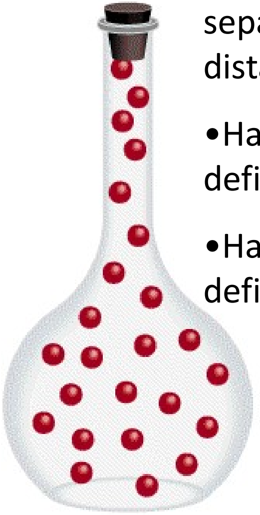


States of matter

Gas

Liquid

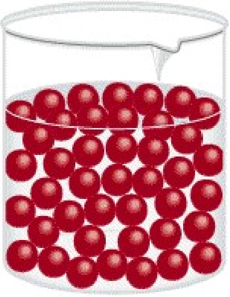
Solid



Gas

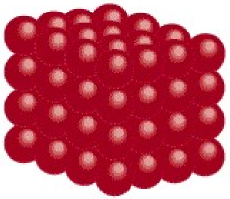
- Molecules are separated by distance
- Have no definite shape
- Have no definite volume

- Molecules are close together but not held so rigidly in position and can move past on another
- Have no definite shape
- Have a definite volume

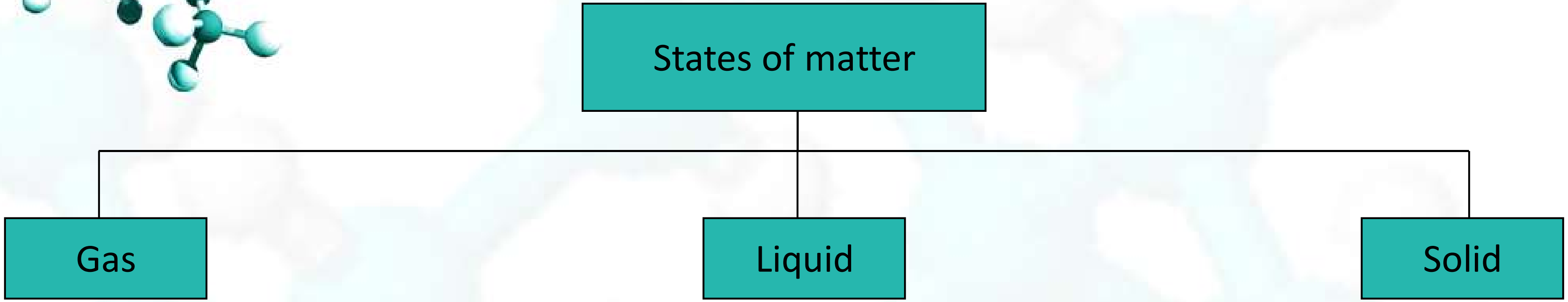


Liquid

- Molecules are held close together in orderly fashion with little freedom of motion.
- Have a definite shape
- Have a definite volume



Solid

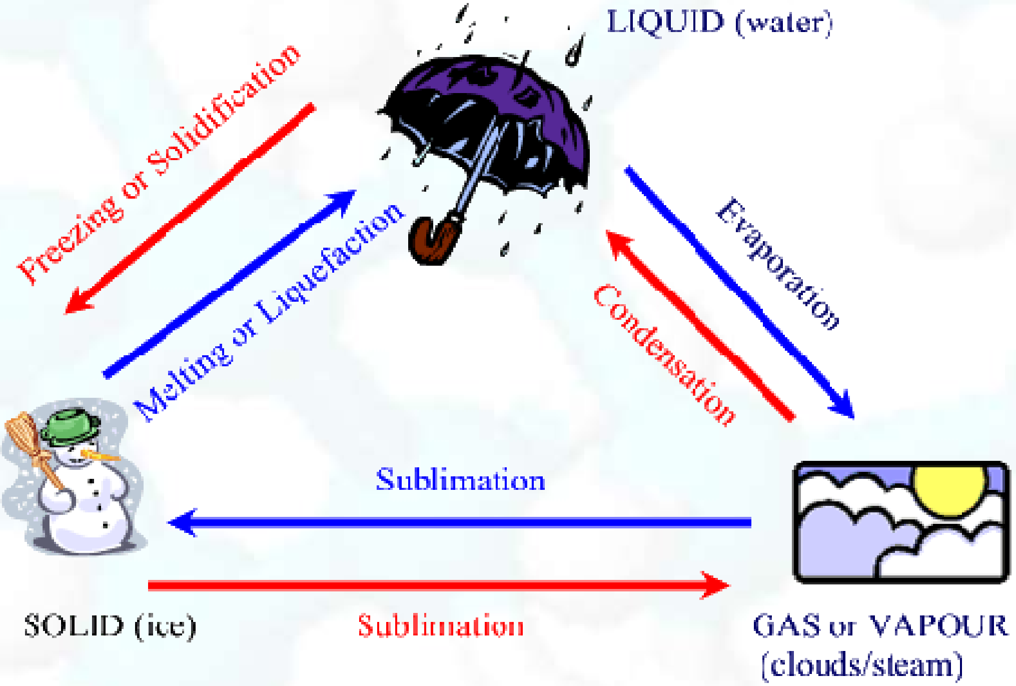


States of matter

Gas

Liquid

Solid



Measurement in chemistry

- In every measurement there is a number followed by a unit.
- Units are essential to stating measurements correctly.
- The international system of units (*SI*) is used world wide to reports result.
- There are seven SI base units as following:

Base Quantity	Name of unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

Measurement in chemistry

- Derived units are unit made up of combination of SI base units such as:

Properties	unit	symbol	Definition
force	Newton	N	Kg m/s^2
Pressure	Pascal	pa	N/m^2 or Kg $/m\ s^2$
Energy	Joule	J	$Kg/m^2\ s^2$ or N m

- SI units are modified in decimal fashion by a series of prefixes.
- A prefix may be added to a unit to produce a multiple of the original unit.

Measurement in chemistry

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000 or 10^{12}	1 terameter (Tm) = 1 x 10^{12} m
giga-	G	1,000,000,000 or 10^{9}	1 gigameter (Gm) = 1 x 10^{9} m
mega-	M	1,000,000 or 10^{6}	1 megameter (Mm) = 1 x 10^{6} m
kilo-	K	1,000 or 10^{3}	1 kilometer (km) = 1 x 10^{3} m
deci-	d	1/10 or 10^{-1}	1 decimeter (dm) = 1 x 10^{-1} m
centi-	c	1/100 or 10^{-2}	1 centimeter (cm) = 1 x 10^{-2}m
milli-	m	1/1,00 or 10^{-3}	1 millimeter (mm) =1 x 10^{-3} m
micro-	µ	1/1,000,000 or 10^{-6}	1 micrometer (µm) = 1 x 10^{-6} m
nano-	n	1/1,000,000,000 or 10^{-9}	1 nanometer (nm) = 1 x 10^{-9} m
pico-	p	1/1,000,000,000,000 or 10^{-12}	1 picometer (pm) =1 x 10^{-12} m

Measurement in chemistry

Unit conversion:

1- prefix to Base unit

eg. Km =====> m

2- Base unit to prefix

eg. m ==========> km

3- prefix to prefix

eg. km ==========>nm

Measurement in chemistry

1- prefix to Base unit

- To remove a prefix from a value, insert the numerical value of the prefix in place of the symbol.

Example : convert 8.53 pm to meters?

Replace p with x 10^{-12}, therefore the answer is 8.53 x 10^{-12} m

2- Base unit to prefix

- To insert a prefix into a value, insert both the prefix and the inverse of its numerical number.

Example: convert 0.000462 g to milligrams ?

0.000462 x 10^{3} mg = 0.462 mg

3- prefix to prefix

Example : convert 6 km to nm?

First convert km to m

Replace k with x 10^{3}, therefore the answer is 6 x 10^{3} m

Then convert m to nm

6 x 10^{3} x 10^{+9} = 6x10^{12} nm

Measurement in chemistry

How many second are in a microseconds ?

(a) 1 x 10^{-1}

[✓ (b) 1 x 10^{-6}]

(c) 1 x 10^{-15}

(d) 1 x 10^{6}

Which of the following is the smallest distance?

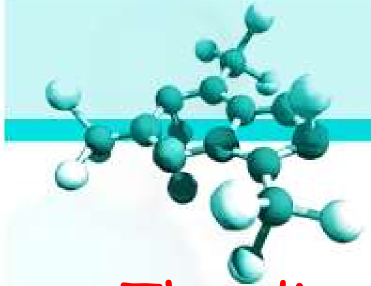
(a) 21 m

(b) 2.1 x 10^{2} cm

(c) 21 mm

[✓ (d) 2.1 x 10^{4} pm]

Put all of them in the same unit



Measurement in chemistry

The diameter of an atom is approximately 1×10^{-7} mm. What is this diameter when expressed in nanometers?

(a) 1×10^{-18} nm
(b) 1×10^{-15} nm
(c) 1×10^{-9} nm
(d) 1×10^{-1} nm

Which of these quantities represents the largest mass?

(a) 2.0×10^{2} mg
(b) 0.0010 kg
(c) 1.0×10^{5} μg
(d) 2.0×10^{2} cg

Put all of them in the same unit

Measurement in chemistry

- There is also a set of non-SI units accepted for use with SI such as litter.
- One common unite of volume is litter (L). A litter is the volume occupied by one cubic decimeter.

1 dm = 10 cm

1 dm^3 = 1000cm^3

1 dm^3= 1 L

1000 cm^3 = 1 L

1 cm^3 = 1mL

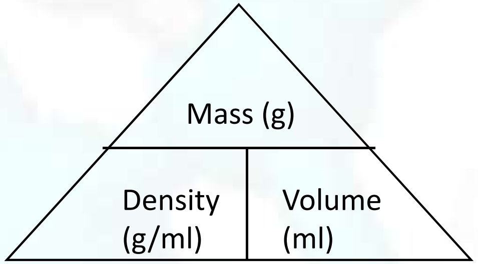
1L=1000 ml
1mL = 10^{-3} L

1mL=1cm^3
1L=1 dm^3

Density

- Density = mass / volume

  d = m/V



Mass (g)

Density (g/ml)

Volume (ml)

- Example:

A student determines that a piece of an unknown material has a mass of 5.854 g and a volume of 7.57 cm^3. What is the density of the material?

d = m/V

= 5.854 (g) / 7.57 (cm^3)

= 0.773 g/cm^3

= 0.773 g/mL

- Example

A piece of silver (Ag) metal weighing 194.3 g is placed in a graduated cylinder containing 242.0 mL of water. The volume of water now reads 260.5 mL. From these data calculate the density of silver?

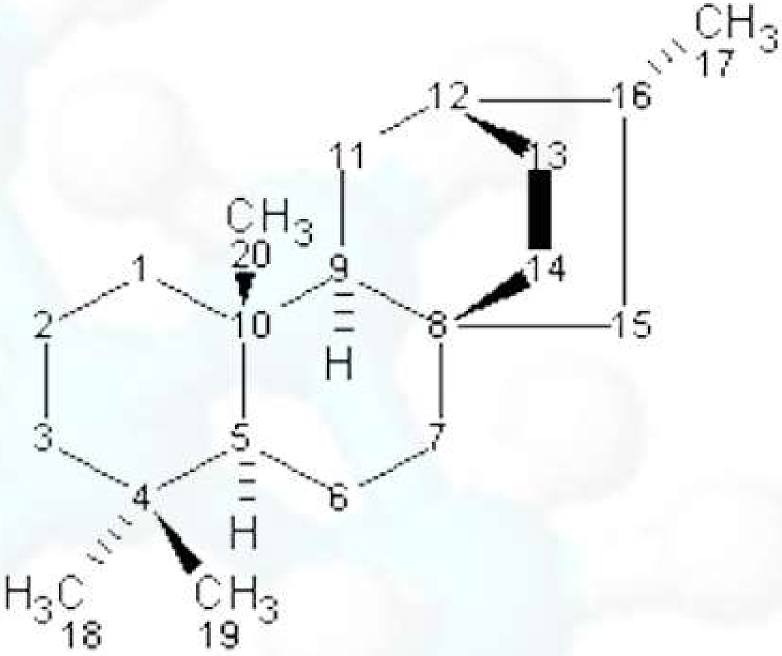
volume of silver = volume of water (after) – volume of water (before) = 260.5 – 242.0 = 18.5 ml

d = m/V

= 194.3 / 18.5

= 10.5 g/ml

Thank you

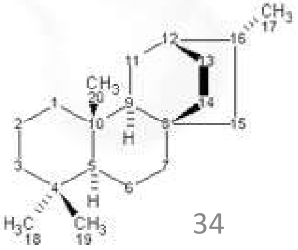


Chapter Two

Atoms, molecules and Ions

- **Chapter 2 Atoms, Molecules and Ions (p42-68)**

- 2.1 The Atomic Theory
- 2.2 The Structure of the Atom
- 2.3 Atomic Number, Mass Number and Isotopes
- 2.4 The Periodic Table
- 2.5 Molecules and Ions
- 2.6 Chemical Formulas
- 2.7 Naming Compounds



The Atomic Theory

- Dalton’s atomic theory has four assumptions:

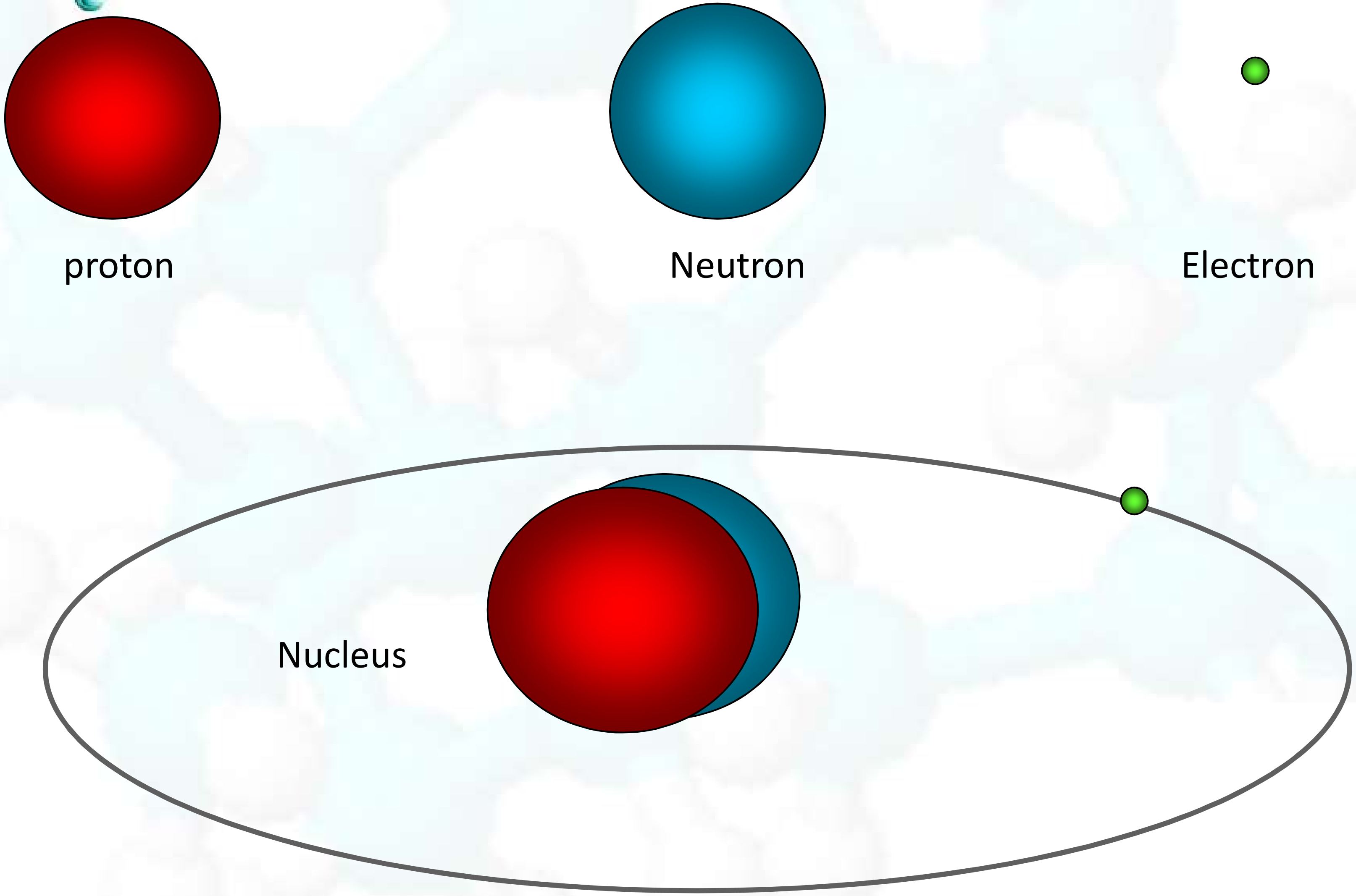
1. Atoms are the building blocks of elements.
2. For the same element all atoms are identical. The atoms for one element are different from the atoms of all other elements.
3. Atoms combine in definite ratio to makes compound.
4. A chemical reaction involves only the separation, combination and rearrangement of atoms. It dos not result in their creation or destruction.

$$H_2O \rightarrow 2H + O$$

What is atom?

Atom is the basic unit of an element that can enter into chemical combination.

Structure of atom



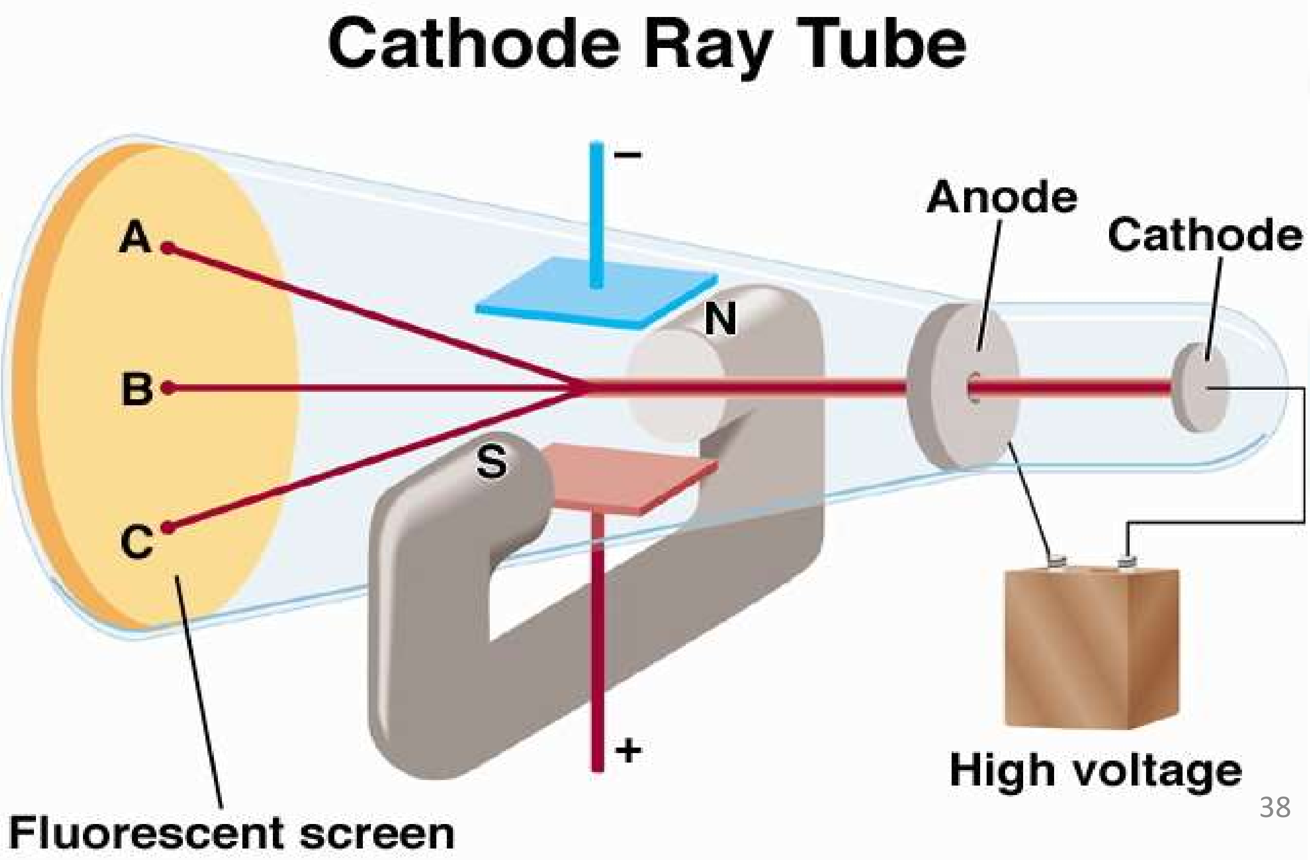
Structure of atom

- Protons and neutrons are together in the *nucleus* of an atom, whereas electrons are in motion in orbits around the central nucleus.
- Protons carry a positive electrical charge, electrons carry a negative charge, and neutrons carry no charge.
- Most atoms are electrically neutral, meaning that they have an equal number of protons and electrons

Particle	Symbol	Charge	Mass
Electron	e^-	-1	0.0005486 amu
Proton	p^+	+1	1.007276 amu
Neutron	n^0	0	1.008665 amu

Structure of atom

- J.J. Thomson is the scientist who determine the ratio between electronic charge to the mass of an electron using a cathode ray tube experiment



Structure of atom

- What is amu?

amu is atomic mass unit

1 amu = $1.66053886 \times 10^{-27}$ kilograms

Atomic Number and Mass Number

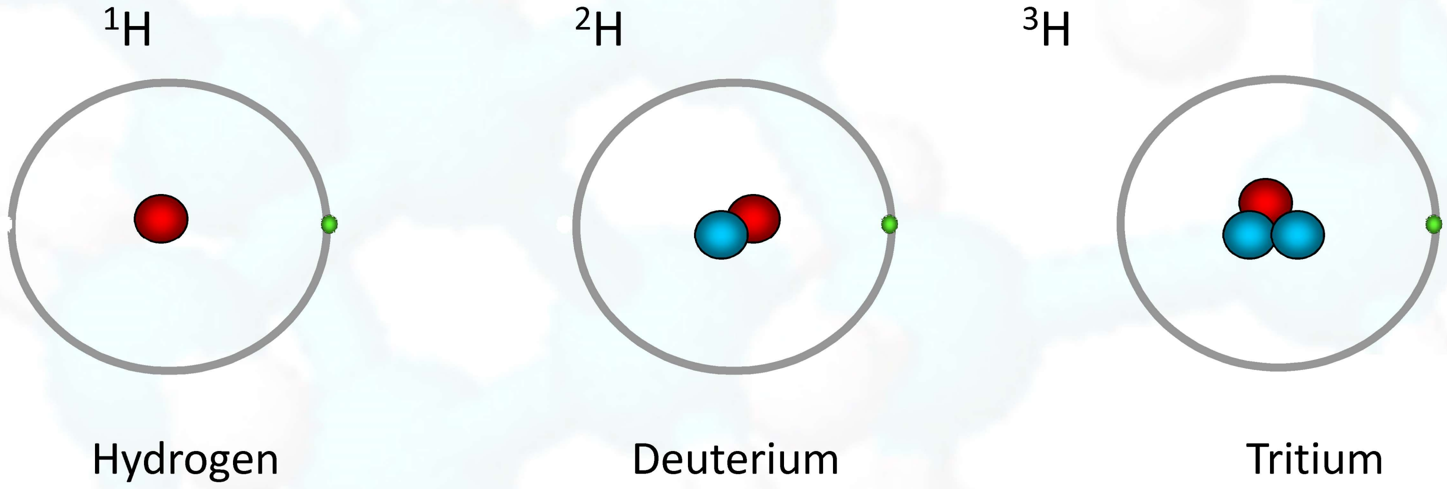
- All atoms can be identified by the number of proton and neutrons they contain.
- ***Atomic number*** is the number of proton in the nucleus.
- In neutral atom the number of proton is equal to the number of electron thus the atomic number also refer to the number of electron.
- The chemical identity of an atom can be determined solely by the atomic number.
- Example: the atomic number of Fluorine is 9 this means that each fluorine atom have 9 protons and 9 electrons.

Any atom have 9 proton is Fluorine.

- ***Mass number*** is the total number of neutrons and protons present in the nucleus of an atom of an element.

Structure of atom

- Thus, number of neutrons = atomic mass - atomic number
- Isotopes are atoms that have the same atomic number (proton number) but different mass numbers.
- This mean the number of neutrons is the only difference between isotopes. For example there are 3 isotopes for hydrogen



Structure of atom

$^{A}_{Z}X$

- Example: Carbon (atomic number 6) has three natural isotopes with atomic weights of 12, 13 and 14.

isotope	#p	#n
======	==	==
C-12	6	6
C-13	6	7
C-14	6	8

- With exception of hydrogen, which has different names for each of its Isotopes, isotopes of element are identified by their mass number. For example carbon isotopes are called : carbon-12 , carbon 13 and carbon 14

Structure of atom

- Important notice : as chemical properties of element determine by the number of proton and electron. Thus isotopes of the same element have similar chemistries, forming the same type of compound and displaying similar reactivity.
- Do you understand Isotopes?
- How many protons, neutrons, and electrons are in $^{14}_{6}C$?

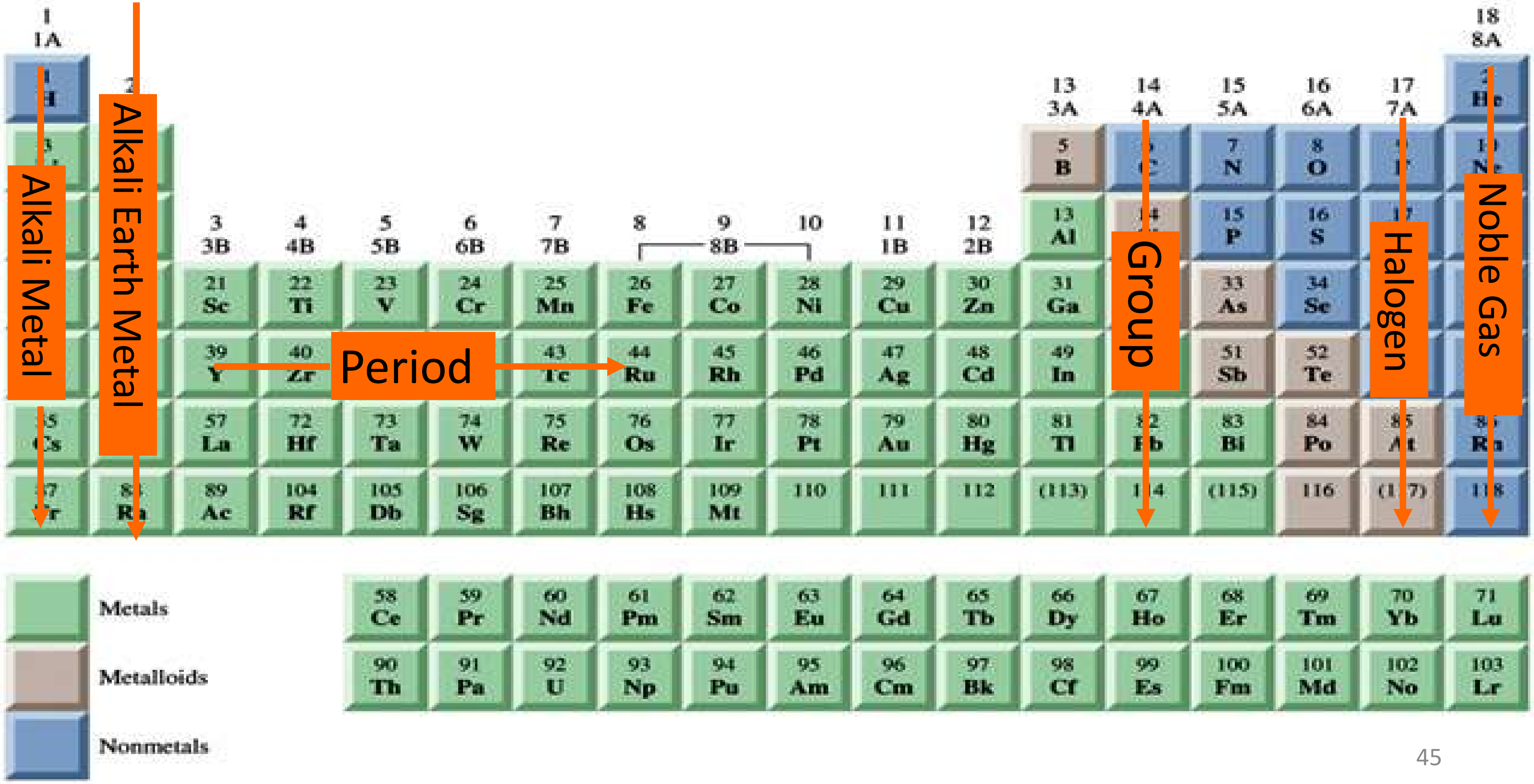
6 protons, 8 (14 - 6) neutrons, 6 electrons

Periodic Table

- Between 1800 and 1900 more than half the elements were discovered.
- Scientist tried to organize all the elements based on similarities which led to the porn of the periodic table.
- The periodic table is a chart in which elements having similar chemical and physical properties are grouped together.

Periodic Table

Modern Periodic Table



1 1A	2	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3	[illegible]											5 B	6 C	7 N	8 O	9 F	10 Ne
[illegible]	[illegible]											13 Al	14	15 P	16 S	17	[illegible]
[illegible]	[illegible]	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	[illegible]	33 As	34 Se	[illegible]	[illegible]
[illegible]	[illegible]	39 Y	40 Zr	[illegible]	[illegible]	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	[illegible]	51 Sb	52 Te	[illegible]	[illegible]
55 Cs	[illegible]	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112	(113)	114	(115)	116	(117)	118

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals

🟫 Metalloids

🟦 Nonmetals

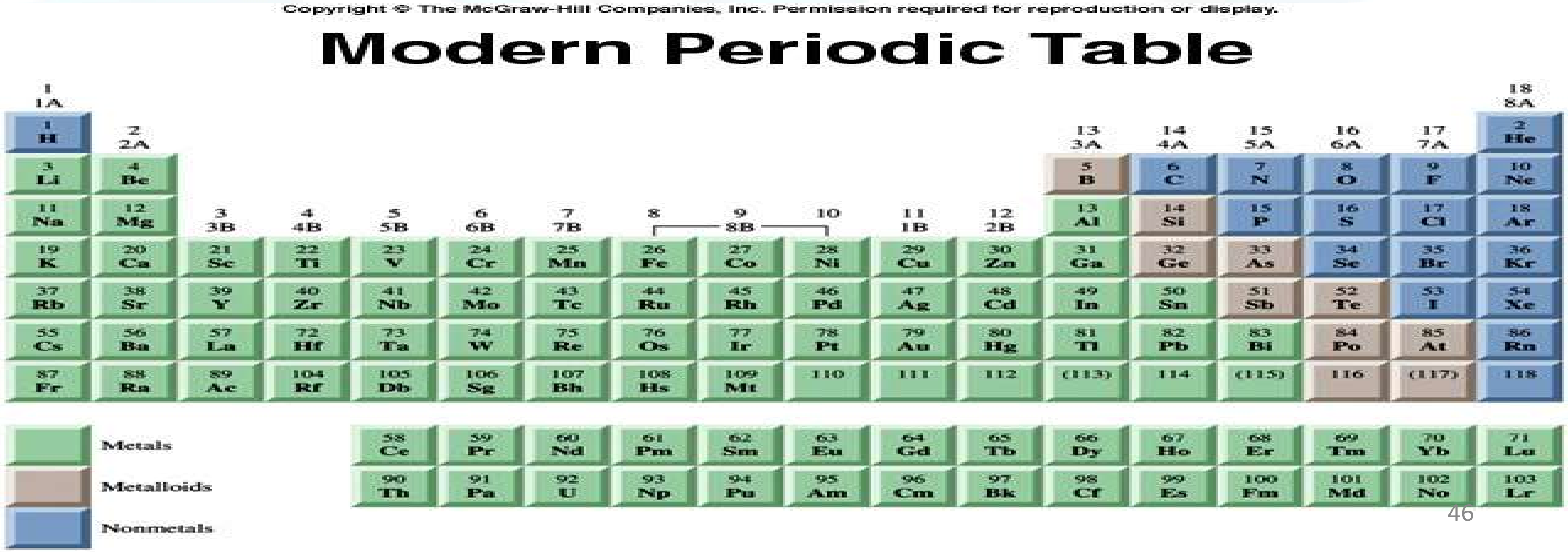
Periodic Table

Elements in the periodic table are divided into three categories:

Metal: (in green colour, Most elements) is a good conductor of heat and electricity

Nonmetal: (in blue colour, 17 elements) is a poor conductor of heat and electricity

Metalloid: (in brown colour, 8 elements) has properties that are intermediate between those of metals and nonmetals



Modern Periodic Table

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112	(113)	114	(115)	116	(117)	118
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals

🟫 Metalloids

🟦 Nonmetals

Molecules

- *Molecules* : is an aggregate of at least two atoms or more in a definite arrangement held together by chemical forces.
- What is the difference between molecule and compound?

Molecules may contain two atoms of the same element or atoms of different elements.

Compound only contain two or more elements

Thus all compounds are molecules but not all molecules are compounds.

Molecules or compounds ?

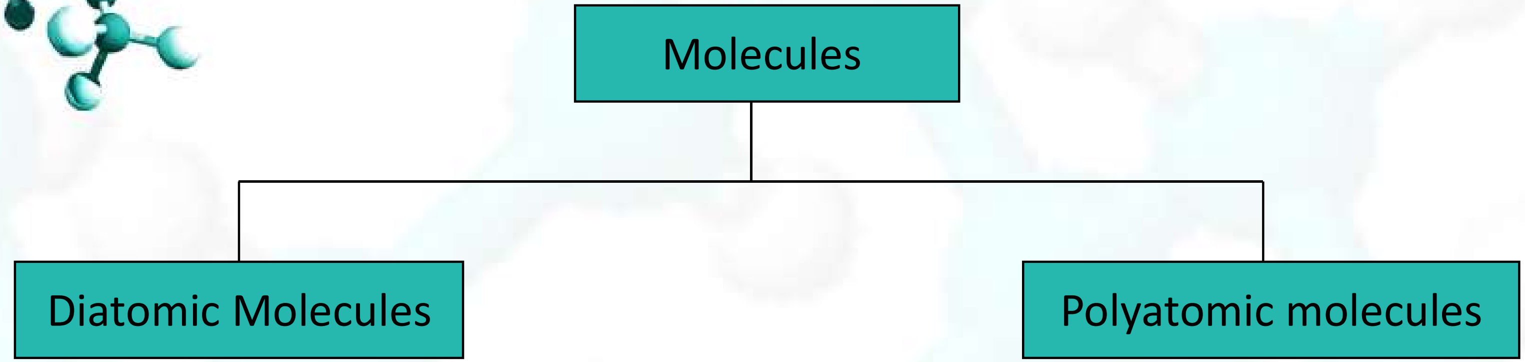
H_2O , I_2 , NaCl , H_2 , O_2 , NaOH

compound

H_2O , NaCl , NaOH

All molecules

Chapter Two / Atoms, molecules and Ions



Molecules

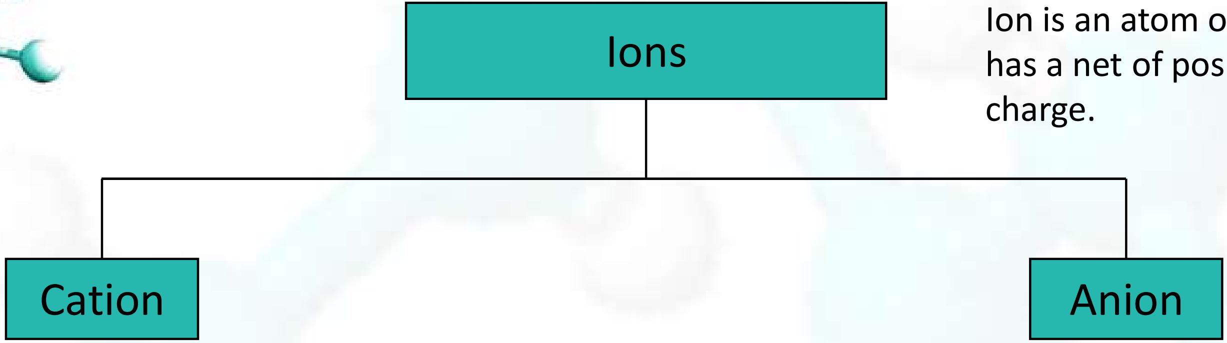
Diatomic Molecules

Polyatomic molecules

Molecules contain only two atoms. Example : H_2 , O_2 N_2 , HCl , CO

Molecules contain more than two atoms. Example : H_2O, NH_3, O_3

Chapter Two / Atoms, molecules and Ions



Ions

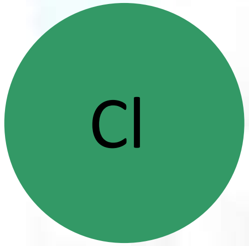
Ion is an atom or a group of atoms that has a net of positive or negative charge.

Cation

An ion with positive charge

If a neutral atom **loses** one or more electrons it becomes a cation.

11 protons
11 electrons



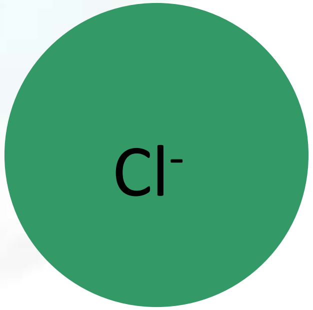
17 protons
17 electrons

Anion

An ion with negative charge

If a neutral atom **gains** one or more electrons it becomes an anion.

11 protons
10 electrons



17 protons
18 electrons

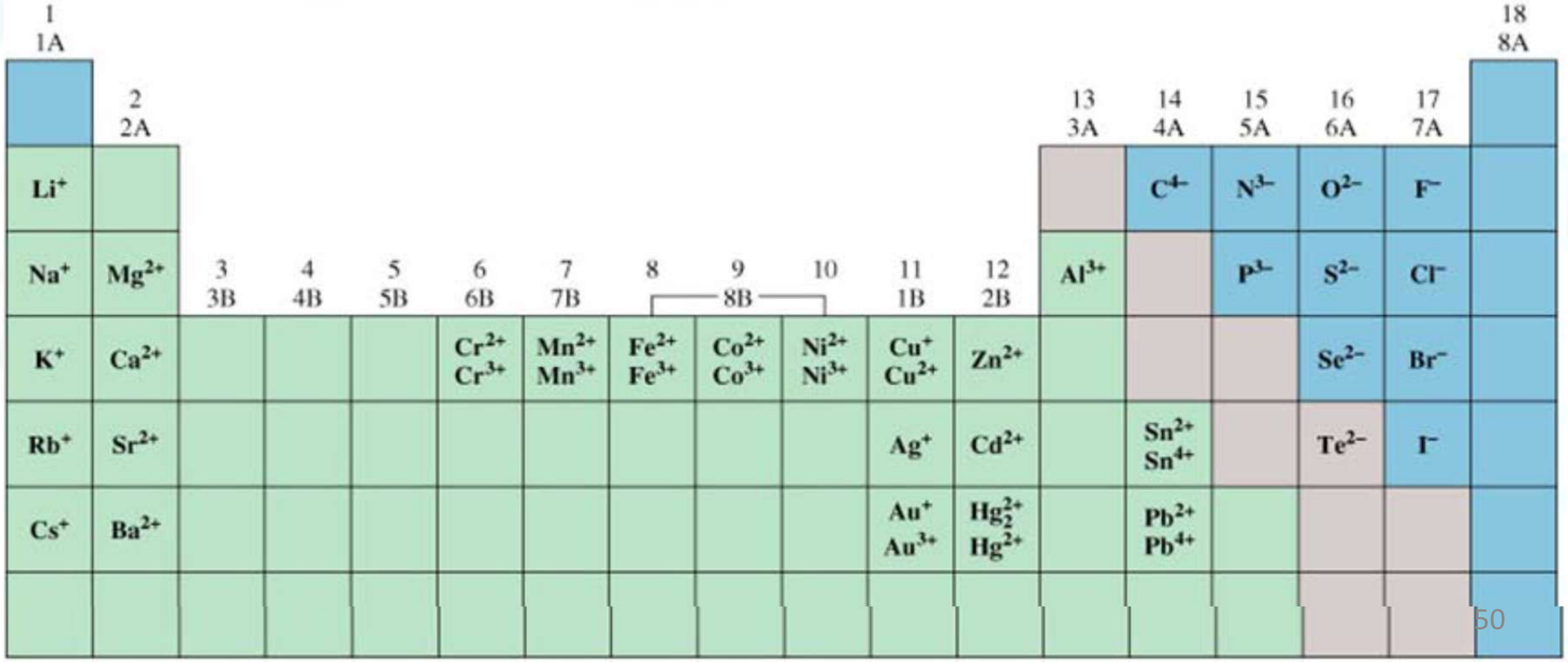
Ions

Monatomic ions

Ions contain only one atom. Example : Na^+ , Mg^{+2} , S^{-2} ,

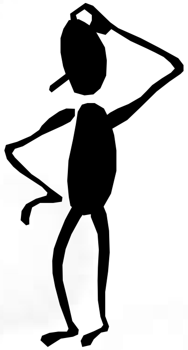
Polyatomic ions

Ions contain more than one atom. Example : NH_4^+ , CN^- , OH^- ,



1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
Li^+													C^{4-}	N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+} Ni^{3+}	Cu^+ Cu^{2+}	Zn^{2+}				Se^{2-}	Br^-	
Rb^+	Sr^{2+}									Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}		Te^{2-}	I^-	
Cs^+	Ba^{2+}									Au^+ Au^{3+}	Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				

Ions



Do You Understand Ions?

How many protons and electrons are in ${}^{27}_{13}\mathrm{Al}^{3+}$?
Proton = 13 , electron = 13-3 = 10
Neutron = 27 – 13 = 14

How many protons and electrons are in ${}^{78}_{34}\mathrm{Se}^{2-}$?
Proton = 34 , electron = 34 +2 = 36
Neutron = 78 – 34 = 44

Chemical Formula

Molecular Formula

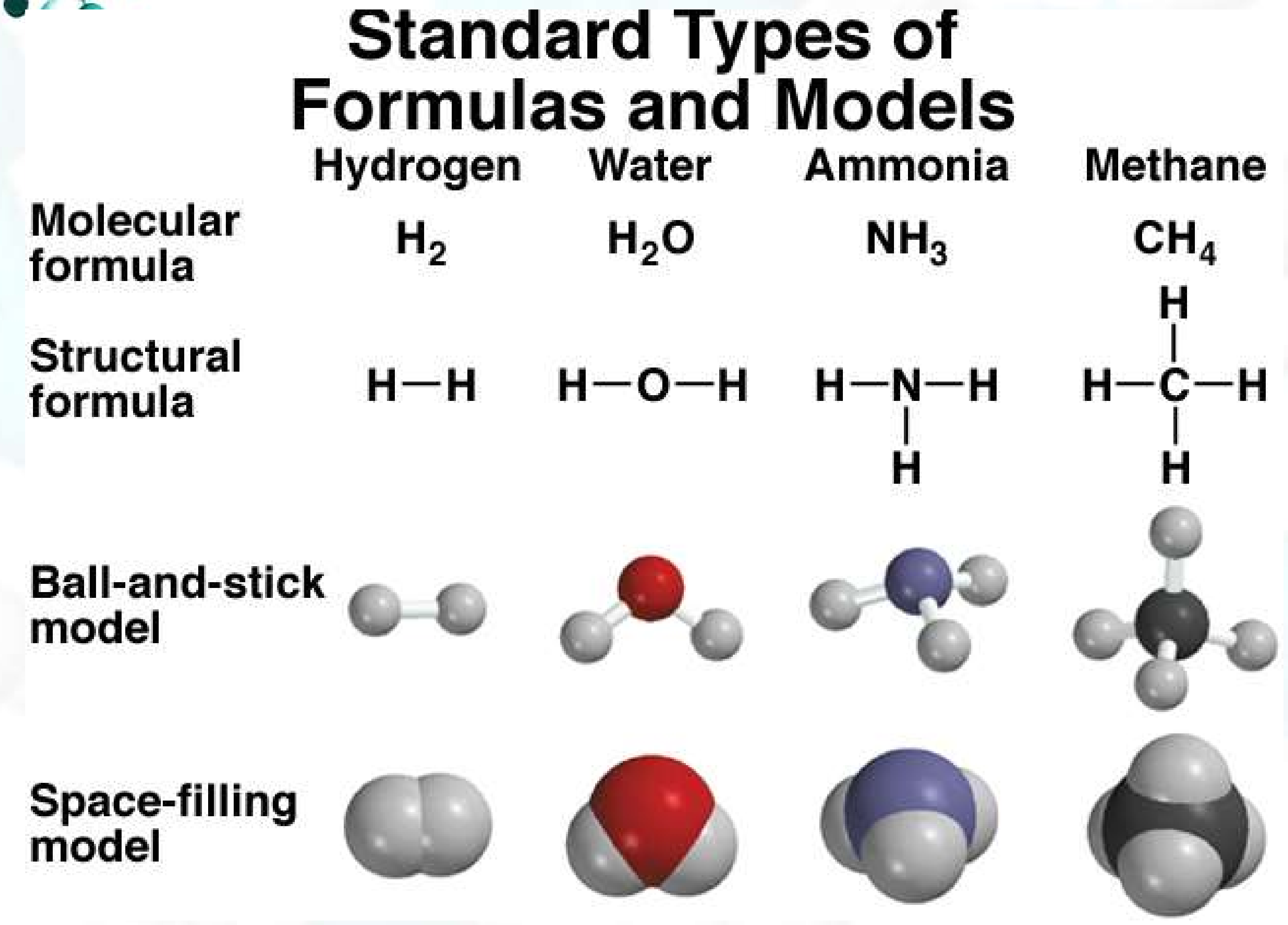
It's a formula shows the exact number of atoms of each elements in the smallest unit of a substance

- $C_6H_{12}O_2$
- N_2H_4
- O_3
- $Cl_2C_2H_4$
- $C_8H_{10}N_4O_4$
- Al_2O_3

Empirical Formula

shows the simplest whole-number ratio of the atoms in a substance

- C_3H_6O
- NH_2
- O
- $ClCH_2$
- $C_4H_5N_2O_2$
- Al_2O_3



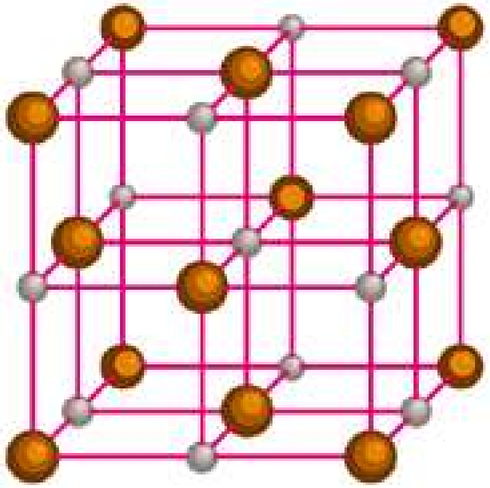
Standard Types of Formulas and Models

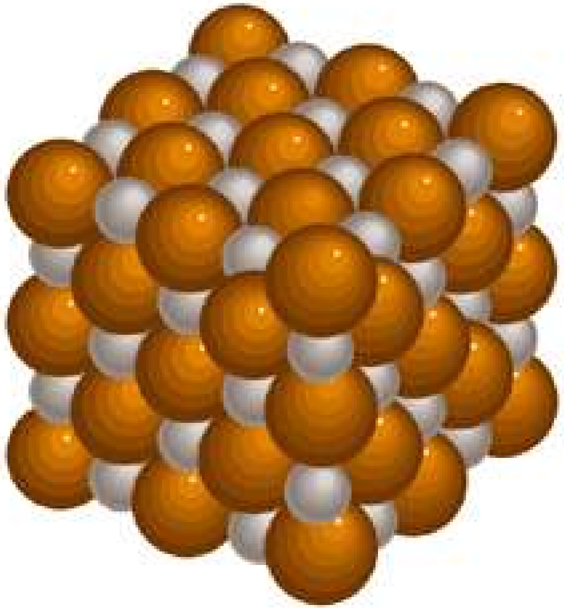
	Hydrogen	Water	Ammonia	Methane
Molecular formula	H_2	H_2O	NH_3	CH_4
Structural formula	$H-H$	$H-O-H$	$H-\underset{\underset{H}{\mid}}{N}-H$	$H-\overset{\overset{H}{\mid}}{\underset{\underset{H}{\mid}}{C}}-H$
Ball-and-stick model				
Space-filling model				

Formula of Ionic Compounds

- Ionic compound consist of a cation and an anion
- the formula is always the same as the empirical formula
- the sum of the charges on the cation and anion in each formula unit must equal zero

The ionic compound NaCl

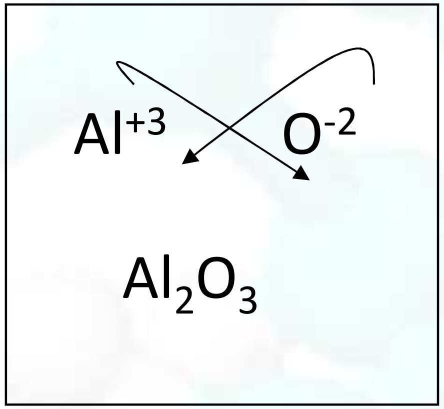




Formula of Ionic Compounds

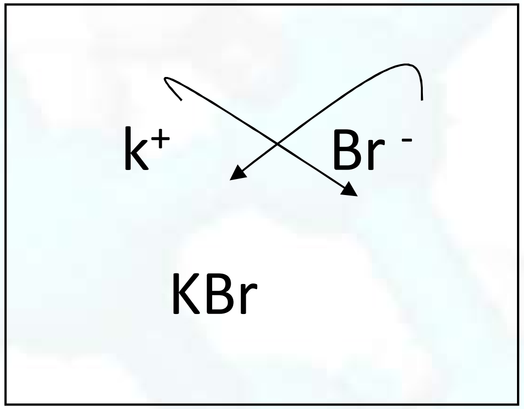
- The subscript of the cation is numerically equal to the charge on the anion, and the subscript on the anion is numerically equal to the charge on the cation

Examples:



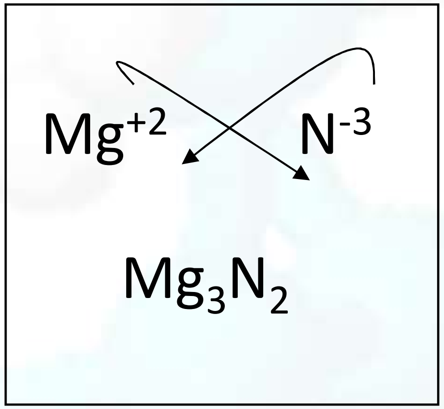
Al^{+3} O^{-2}

Al_2O_3



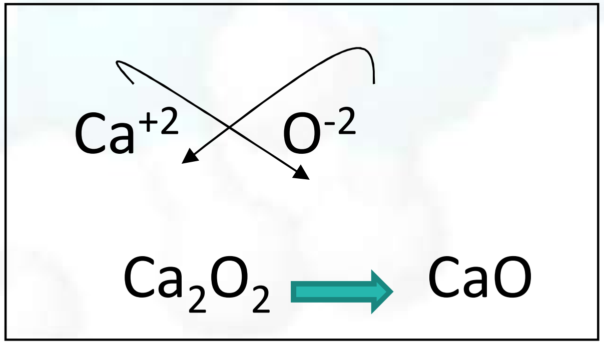
k^{+} Br^{-}

KBr



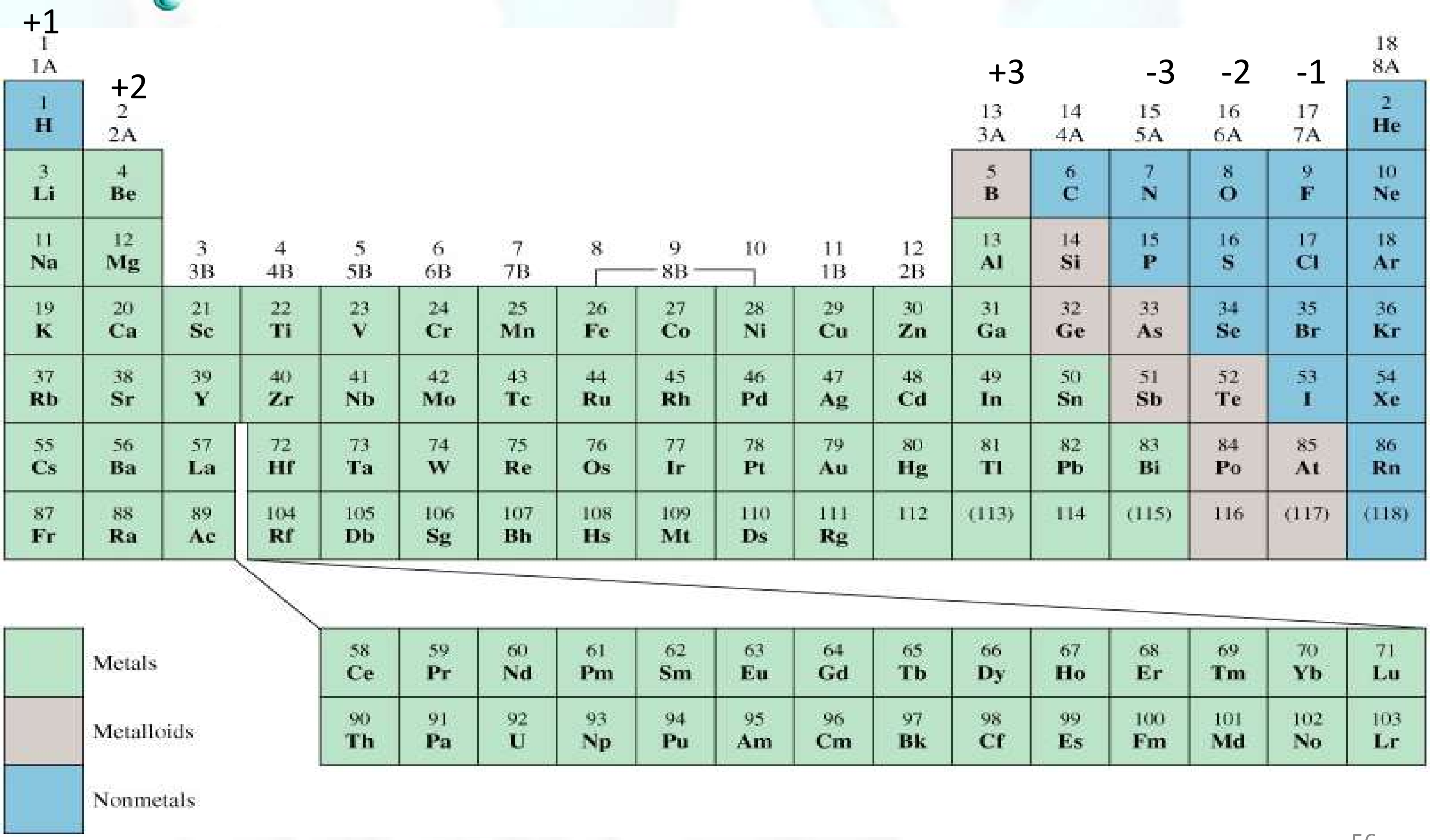
Mg^{+2} N^{-3}

Mg_3N_2



Ca^{+2} O^{-2}

$Ca_2O_2 \rightarrow CaO$



+1 1 1A	+2 2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	+3 13 3A	14 4A	-3 15 5A	-2 16 6A	-1 17 7A	18 8A
1 **H**																	2 **He**
3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
55 **Cs**	56 **Ba**	57 **La**	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
87 **Fr**	88 **Ra**	89 **Ac**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112	(113)	114	(115)	116	(117)	(118)

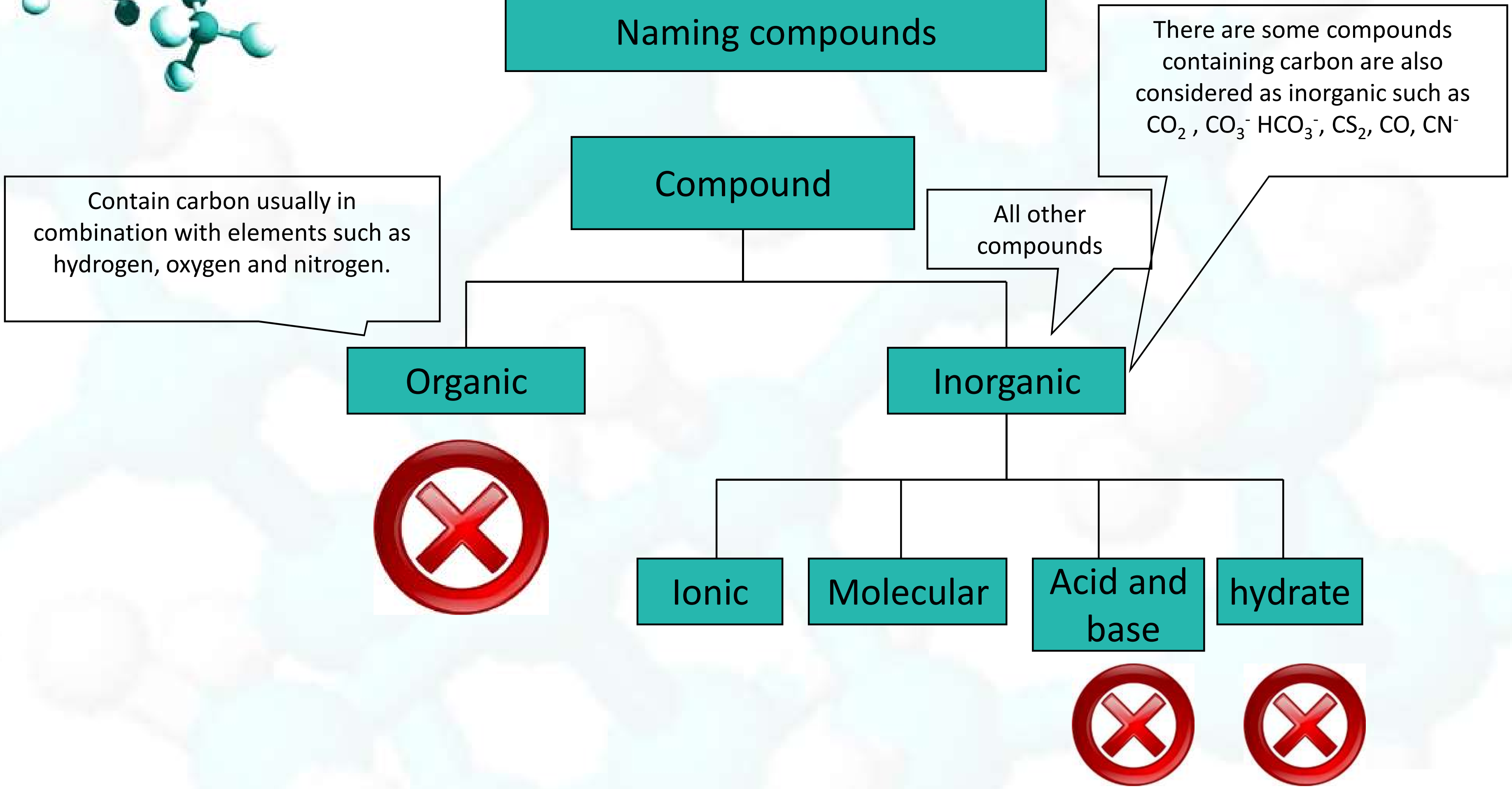
58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

🟩 Metals

▨ Metalloids

🟦 Nonmetals

Naming compounds



Compound

Organic

Contain carbon usually in combination with elements such as hydrogen, oxygen and nitrogen.

Inorganic

All other compounds

There are some compounds containing carbon are also considered as inorganic such as CO_2 , CO_3^- HCO_3^-, CS_2, CO, CN^-

Ionic

Molecular

Acid and base

hydrate

Ionic Compounds

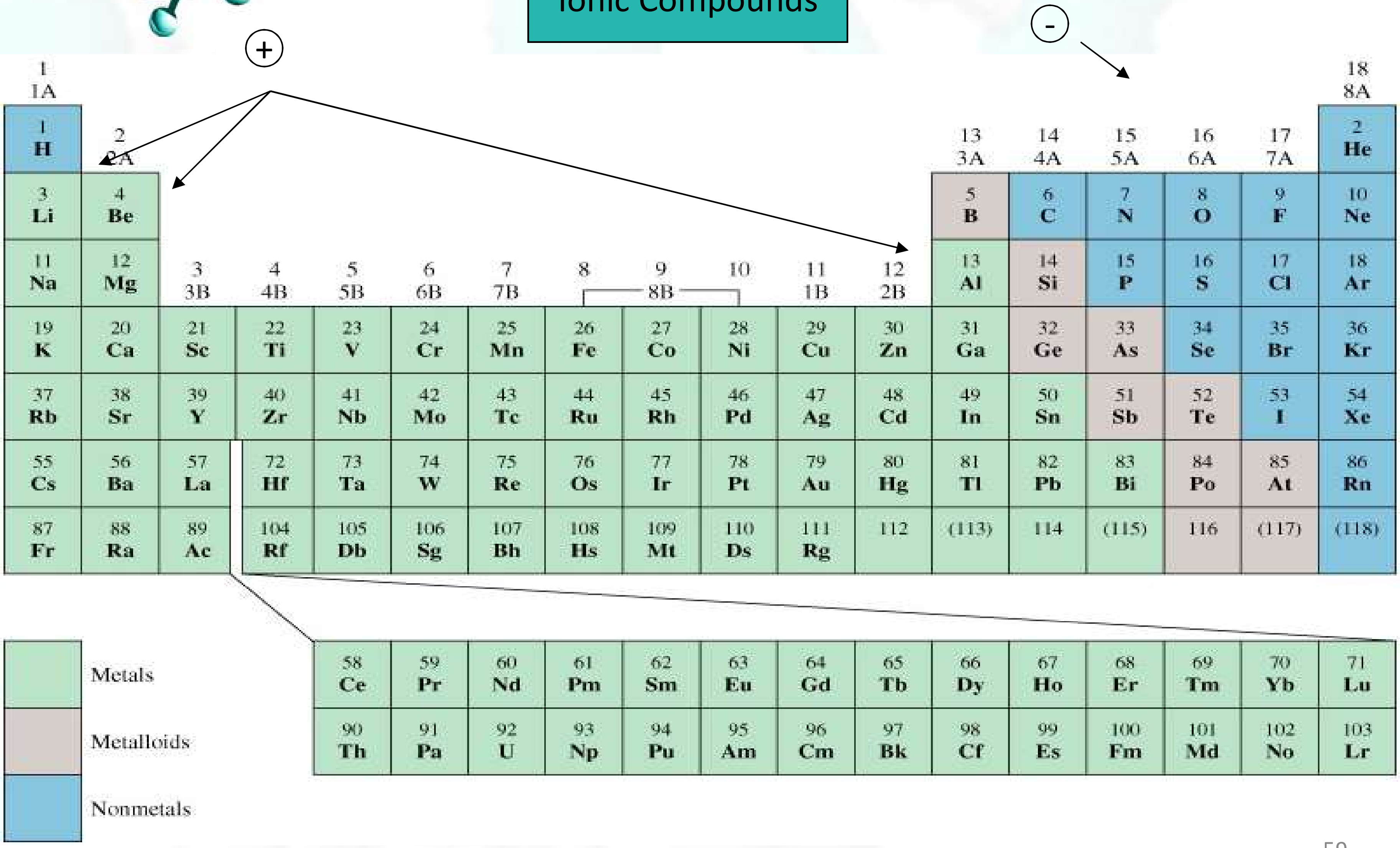
- Mostly metals cation + nonmetal anion
- Rule: Name the metal first, then the nonmetal as -ide.

Example :

NaCl	Sodium chloride
ZnI_2	Zinc iodide
Al_2O_3	Aluminum oxide
Na_3N	Sodium nitride
KBr	Potassium bromide
CaO	Calcium oxide
MgS	Magnesium sulfide

NH_4^+ is called ammonium

Ionic Compounds



1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
🟦 1 H																	🟦 2 He
🟩 3 Li	🟩 4 Be											🟫 5 B	🟦 6 C	🟦 7 N	🟦 8 O	🟦 9 F	🟦 10 Ne
🟩 11 Na	🟩 12 Mg											🟩 13 Al	🟫 14 Si	🟦 15 P	🟦 16 S	🟦 17 Cl	🟦 18 Ar
🟩 19 K	🟩 20 Ca	🟩 21 Sc	🟩 22 Ti	🟩 23 V	🟩 24 Cr	🟩 25 Mn	🟩 26 Fe	🟩 27 Co	🟩 28 Ni	🟩 29 Cu	🟩 30 Zn	🟩 31 Ga	🟫 32 Ge	🟫 33 As	🟦 34 Se	🟦 35 Br	🟦 36 Kr
🟩 37 Rb	🟩 38 Sr	🟩 39 Y	🟩 40 Zr	🟩 41 Nb	🟩 42 Mo	🟩 43 Tc	🟩 44 Ru	🟩 45 Rh	🟩 46 Pd	🟩 47 Ag	🟩 48 Cd	🟩 49 In	🟩 50 Sn	🟫 51 Sb	🟫 52 Te	🟦 53 I	🟦 54 Xe
🟩 55 Cs	🟩 56 Ba	🟩 57 La	🟩 72 Hf	🟩 73 Ta	🟩 74 W	🟩 75 Re	🟩 76 Os	🟩 77 Ir	🟩 78 Pt	🟩 79 Au	🟩 80 Hg	🟩 81 Tl	🟩 82 Pb	🟩 83 Bi	🟫 84 Po	🟫 85 At	🟦 86 Rn
🟩 87 Fr	🟩 88 Ra	🟩 89 Ac	🟩 104 Rf	🟩 105 Db	🟩 106 Sg	🟩 107 Bh	🟩 108 Hs	🟩 109 Mt	🟩 110 Ds	🟩 111 Rg	🟩 112	🟩 (113)	🟩 114	🟩 (115)	🟫 116	🟫 (117)	🟦 (118)

🟩 58 Ce	🟩 59 Pr	🟩 60 Nd	🟩 61 Pm	🟩 62 Sm	🟩 63 Eu	🟩 64 Gd	🟩 65 Tb	🟩 66 Dy	🟩 67 Ho	🟩 68 Er	🟩 69 Tm	🟩 70 Yb	🟩 71 Lu
🟩 90 Th	🟩 91 Pa	🟩 92 U	🟩 93 Np	🟩 94 Pu	🟩 95 Am	🟩 96 Cm	🟩 97 Bk	🟩 98 Cf	🟩 99 Es	🟩 100 Fm	🟩 101 Md	🟩 102 No	🟩 103 Lr

🟩	Metals
🟫	Metalloids
🟦	Nonmetals

Ionic Compounds

- Metal cation can be devided into two catagrories :

1- metal form one type of cation. Example: Alkali group and alkali earth group.

2- metal form more than one type of cation. Example: transient metals. (Fe^{+2}, Fe^{+3}).

- Rule Exceptions:

If have a variable charged metal, then give its charge in the middle with a Roman Numeral in parenthesis. This called stock system.

Example:

$FeCl_2$= Iron (II) Chloride.

$FeCl_3$= Iron (III) chloride

Ionic Compounds

- Nonmetal anion can be divided into two groups:

1- monoatomic anion. Example : Cl^- chloride , Br^- bromide, O^{-2} Oxide.

2- Polyatomic anion. Example : OH^- hydroxide , CN^- cyanide.

- Common names for some anions:

CO_3^{-2} Carbonate, PO_4^{-3} Phosphate, SO_4^{-2} Sulphate

HCO_3^- Bicarbonate, NO_3^- Nitrate, SO_3^{-2} Sulphite

Ionic Compounds

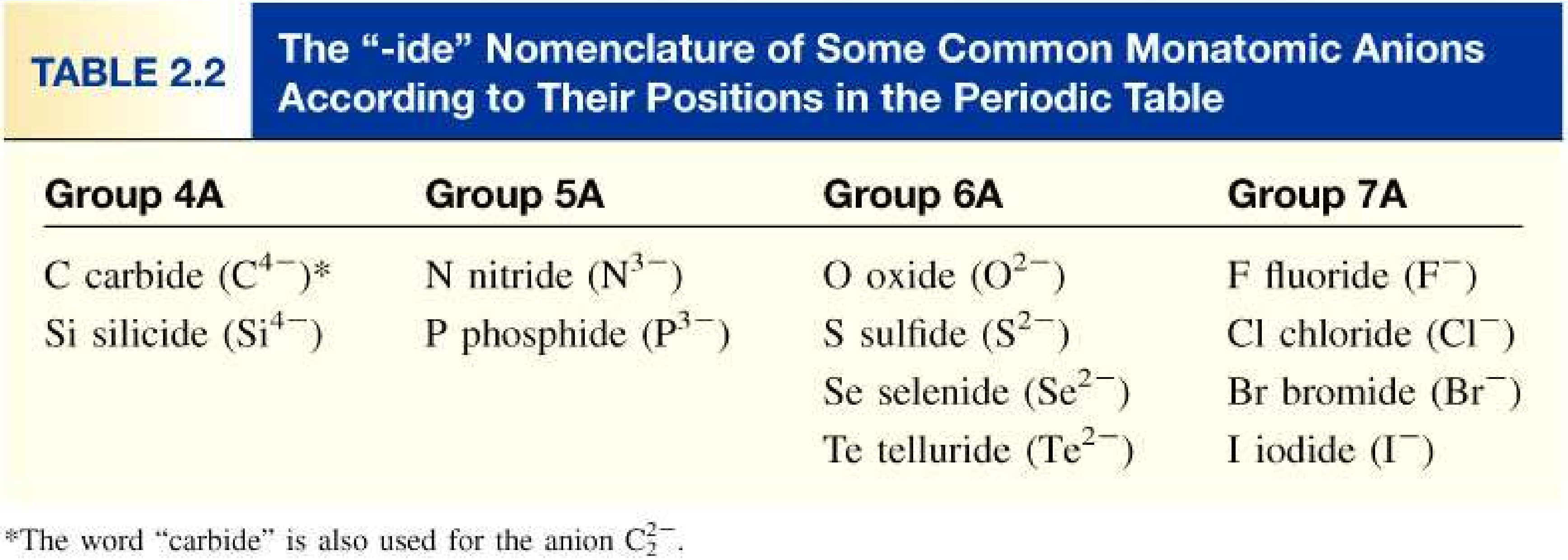


TABLE 2.2 The "-ide" Nomenclature of Some Common Monatomic Anions According to Their Positions in the Periodic Table

Group 4A	Group 5A	Group 6A	Group 7A
C carbide (C^{4-})*	N nitride (N^{3-})	O oxide (O^{2-})	F fluoride (F^{-})
Si silicide (Si^{4-})	P phosphide (P^{3-})	S sulfide (S^{2-})	Cl chloride (Cl^{-})
		Se selenide (Se^{2-})	Br bromide (Br^{-})
		Te telluride (Te^{2-})	I iodide (I^{-})

*The word "carbide" is also used for the anion C_2^{2-}.

Chapter Two / Atoms, molecules and Ions

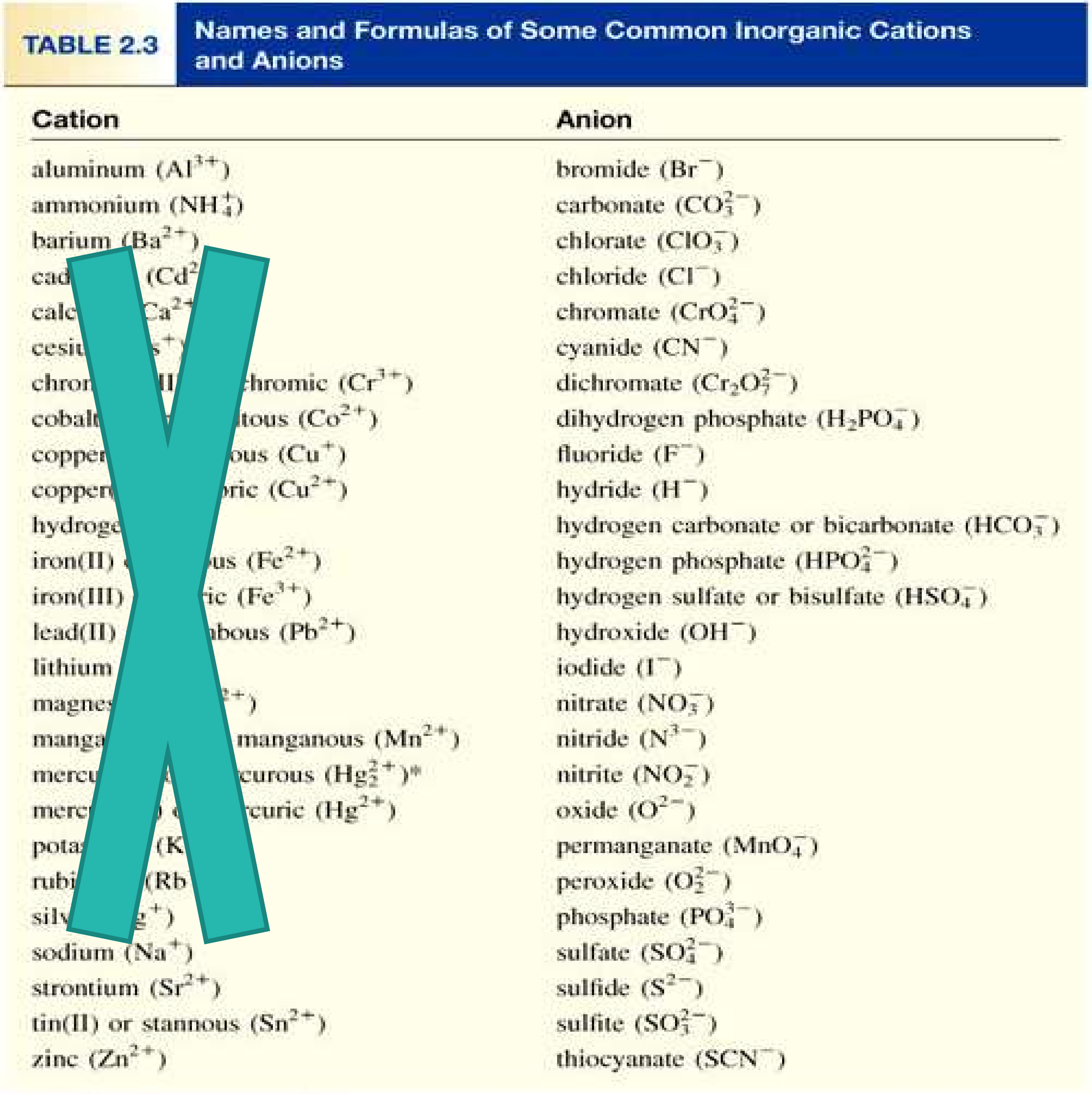


TABLE 2.3 Names and Formulas of Some Common Inorganic Cations and Anions

Cation	Anion
aluminum (Al^{3+})	bromide (Br^-)
ammonium (NH_4^+)	carbonate (CO_3^{2-})
barium (Ba^{2+})	chlorate (ClO_3^-)
cad[illegible] (Cd[illegible]	chloride (Cl^-)
calc[illegible]Ca[illegible]	chromate (CrO_4^{2-})
cesiu[illegible]s^{+})	cyanide (CN^-)
chrom[illegible]II[illegible]hromic (Cr^{3+})	dichromate ($Cr_2O_7^{2-}$)
cobalt[illegible]ltous (Co^{2+})	dihydrogen phosphate ($H_2PO_4^-$)
copper[illegible]ous (Cu^+)	fluoride (F^-)
copper[illegible]ric (Cu^{2+})	hydride (H^-)
hydroge[illegible]	hydrogen carbonate or bicarbonate (HCO_3^-)
iron(II) [illegible]us (Fe^{2+})	hydrogen phosphate (HPO_4^{2-})
iron(III) [illegible]ric (Fe^{3+})	hydrogen sulfate or bisulfate (HSO_4^-)
lead(II) [illegible]bous (Pb^{2+})	hydroxide (OH^-)
lithium [illegible]	iodide (I^-)
magne[illegible]$^{2+}$)	nitrate (NO_3^-)
manga[illegible] manganous (Mn^{2+})	nitride (N^{3-})
mercu[illegible]curous (Hg_2^{2+})*	nitrite (NO_2^-)
merc[illegible]curic (Hg^{2+})	oxide (O^{2-})
pota[illegible] (K[illegible]	permanganate (MnO_4^-)
rubi[illegible](Rb[illegible]	peroxide (O_2^{2-})
silv[illegible]g^{+})	phosphate (PO_4^{3-})
sodium (Na^+)	sulfate (SO_4^{2-})
strontium (Sr^{2+})	sulfide (S^{2-})
tin(II) or stannous (Sn^{2+})	sulfite (SO_3^{2-})
zinc (Zn^{2+})	thiocyanate (SCN^-)

Ionic Compounds

- Example 1 :

Name the following compounds :

a) $Cu(NO_3)_2$ b) KH_2PO_4 C) NH_4ClO_3

b) Cu from transition metal then it have more than one form of cation, NO_3^{-1} has common name nitrate thus

copper (II) nitrate

b) Potassium from group 1A thus it form only one type of cation, $H_2PO_4^{-1}$ has common name dihydrogen phosphate thus

Potassium dihydrogen phosphate.

C) NH_4^{+1} has common name ammonium, ClO_3^{-1} has common name chlorate thus

Ammonium chlorate

Ionic Compounds

- Example 2:

Write the chemical formula of the following compounds :

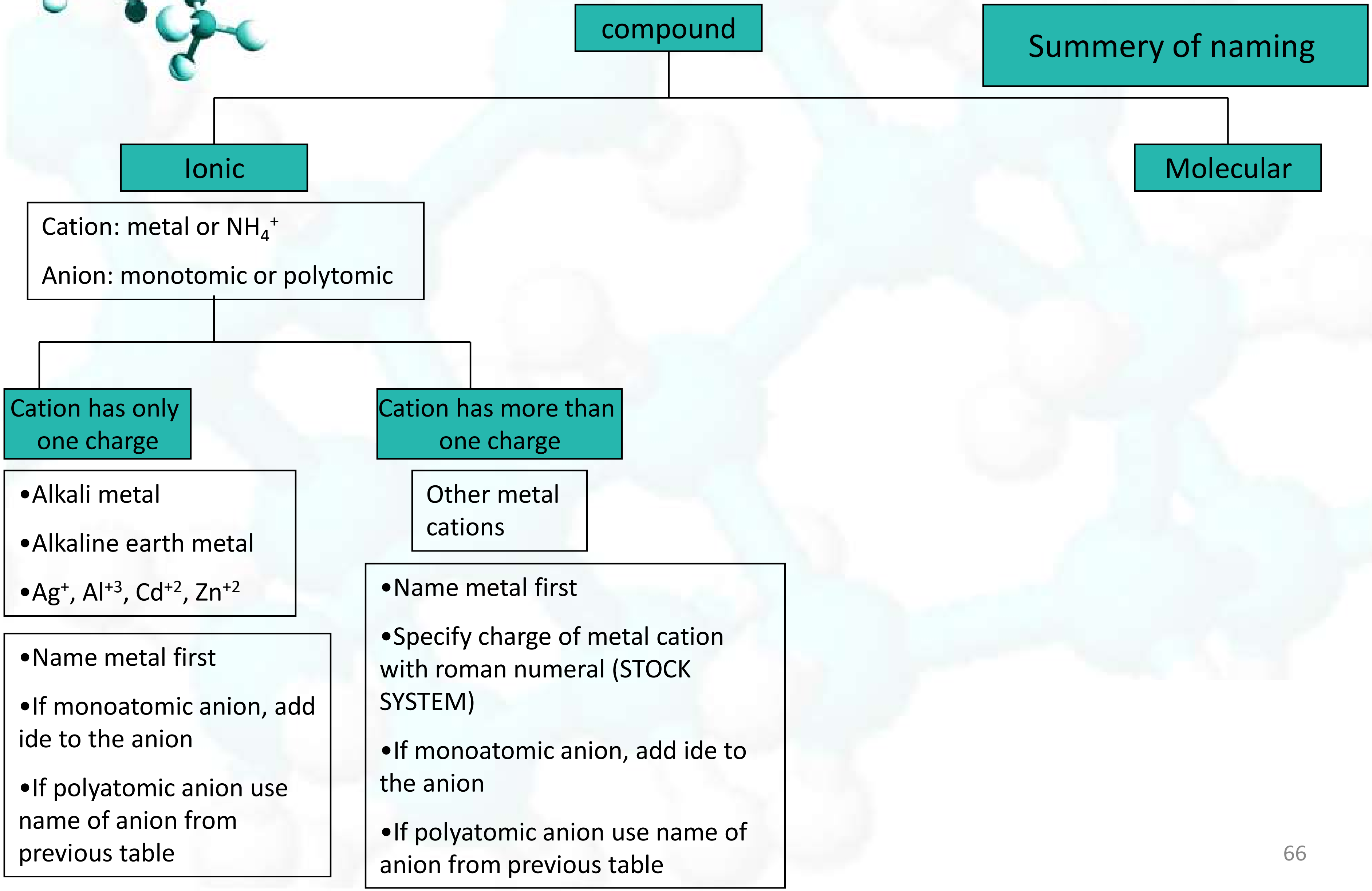
a) Mercury(II) nitrite, b) cesium sulfide c) calcium phosphate.

a) Hg^{+2} NO_2^{-1} ===== $Hg(NO_2)_2$

b) Cs^{+1} S^{-2} =======Cs_2S

c) Ca^{+2} PO_4^{-3} ========$Ca_3(PO_4)_2$

Summery of naming



compound

Ionic

Molecular

Cation: metal or NH_4^+

Anion: monotomic or polytomic

Cation has only one charge

- Alkali metal
- Alkaline earth metal
- Ag^+, Al^{+3}, Cd^{+2}, Zn^{+2}

- Name metal first
- If monoatomic anion, add ide to the anion
- If polyatomic anion use name of anion from previous table

Cation has more than one charge

Other metal cations

- Name metal first
- Specify charge of metal cation with roman numeral (STOCK SYSTEM)
- If monoatomic anion, add ide to the anion
- If polyatomic anion use name of anion from previous table

Molecular Compounds

- Mostly nonmetal + nonmetal or

  nonmetal + metalloid

Rule: Name the first element, then the second element as -ide.

Example:

HCl Hydrogen chloride

HBr Hydrogen bromide

SiC silicon carbide

- Its common that one pair of element can form different compounds. Example:

CO	SO_2	NO_2
CO_2	SO_3	N_2O_4

Molecular Compounds

CO	carbon	monoxide
CO_2	carbon	dioxide
SO_2	sulfur	dioxide
SO_3	sulfur	trioxide
N_2O_4	dinitorgen	tetraoxide

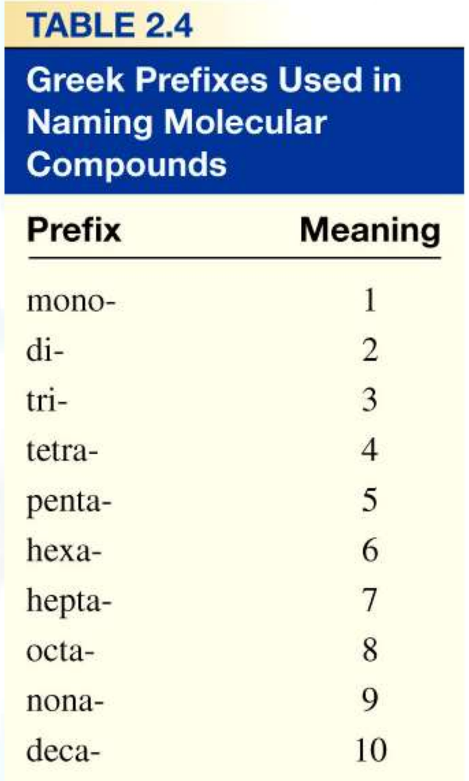


TABLE 2.4

Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

- Name as before and add the number of atom before each elements.
- Note that for the first element the prefix mono can be omitted.
- Compounds containing hydrogen are exception from this rule. Example:

B_2H_6 diborane, CH_4 methane, NH_3 ammonia , SiH_4 silane, PH_3 phosphine
H_2O water, H_2S hydrogen sulfide

Molecular Compounds

- Example1:

Name the following compounds:

a) PCl_3 b) CCl_4 c) P_2Cl_5

a) Phosphorus trichloride

b) Carbon tetrachloride

c) diphosphorus pentachloride

Molecular Compounds

- Example2:

Write the chemical formula of the following compounds:

a) Carbon disulfide b) disilicon hexabromide c) sulfur tetrafluoride

a) CS_2

b)Si_2Br_6

c) SF_4

Summery of naming

compound

Ionic

Cation: metal or NH_4^+

Anion: monotomic or polytomic

- **Cation has only one charge**
  - Alkali metal
  - Alkaline earth metal
  - Ag^+, Al^{+3}, Cd^{+2}, Zn^{+2}

  - Name metal first
  - If monoatomic anion, add ide to the anion
  - If polyatomic anion use name of anion from previous table
- **Cation has more than one charge**
  - Other metal cations

  - Name metal first
  - Specify charge of metal cation with roman numeral (STOCK SYSTEM)
  - If monoatomic anion, add ide to the anion
  - If polyatomic anion use name of anion from previous table

Molecular

Nonmetal + nonmetal

Nonmetal + metalloid

- **Pair Form one type of compound**
  - Name first element
  - add ide to the name of second element
- **Pair Form more than one type of compound**
  - Name first element
  - add ide to the name of second element
  - Add the prefix (prefix mono usually omitted for the first element

homework

- Name these compounds:

a) Na_2CrO_4

b) K_2HPO_4

c) HBr

d) Li_2CO_3

e) $K_2Cr_2O_7$

f) NH_4NO_2

g) PF_3

h) PF_5

i) P_4O_6

j) CdI_2

k) $SrSO_4$

l) $Al(OH)_3$

m) Na_2CO_3

n) $FeCl_2$

o) FeO

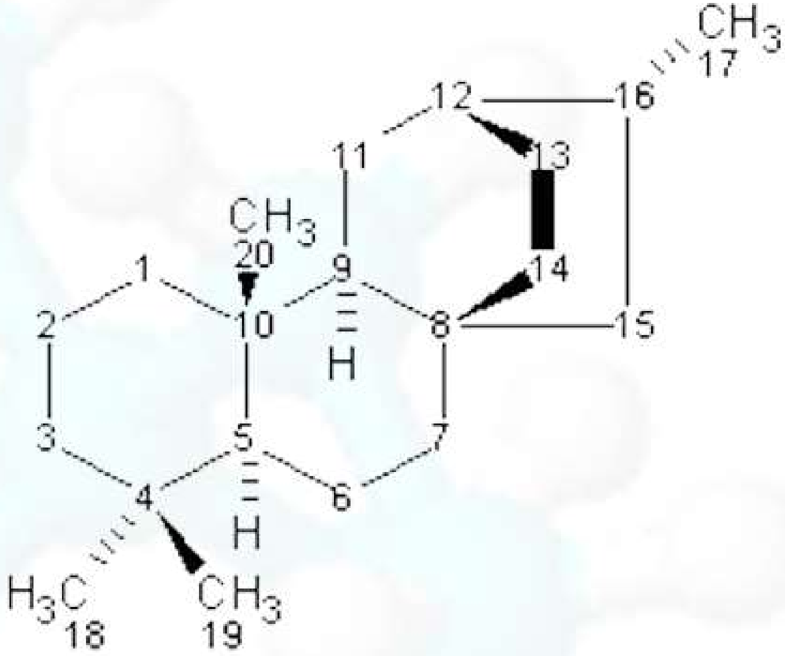
p) Fe_2O_3

homework

- Write the chemical formula for the following compound:

a) Rubidium nitrite
b) potassium sulphide
c) magnesium phosphate
d) calcium hydrogen phosphate
e) potassium dihydrogen phosphate
f) iodine heptaflouride
g) ammonium sulphate
h) silver perchlorate
i) boron trichloride
j) copper (I) cyanide
k) copper (II) cyanide
l) lead (II) carbonate
m) lead (IV) carbonate

Thank you

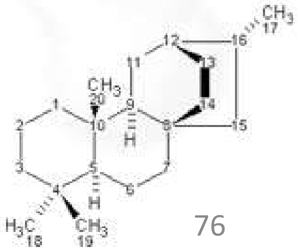


Chapter Three

Mass Relationships in Chemical Reactions

Chapter 3 Mass Relationships in Chemical Reaction (p78-107)

- 3.1 Atomic Mass
- 3.2 Avogadro’s Number and Molar Mass of an element
- 3.3 Molecular Mass
- 3.5 Percent Composition of Compounds
- 3.6 Determining the Formula of a Compound: Empirical, Molecular and Structural Formulas
- 3.7 Chemical Equations and Reactions, Balancing Chemical Equations
- 3.8 Stoichiometric Calculations: Amount of Reactants and Products
- 3.9 Calculation Involving Limiting Reagents
- 3.10 Reaction Yield

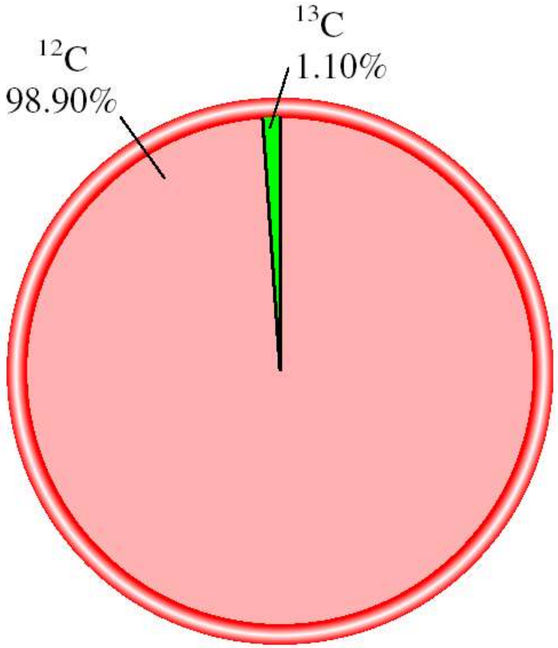


Atomic mass

- Atomic mass = protons + neutrons
- The atom is too small to be weighted.
- However, we can determine the mass of one atom relative to another.
- Atomic mass = the mass of atom in amu
- amu is the mass that exactly equal to one-twelfth the mass of one carbon-12 (^{12}C) atom.
- By definition: 1 atom ^{12}C weight 12 amu.
- Setting the atomic mass of carbon-12 at 12 amu provied the standard for measuring the atomic mass for the other elements.

Average atomic mass

- Most element in nature have more than one isotopes. This mean when calculating the atomic mass we should calculated for all isotopes then take the average this called *Average atomic mass.*
- Natural abundance is the abundance of isotopes in nature.



Average atomic mass

- Average atomic mass= sum of (natural abundance x atomic mass) for each isotope

Example: calculate the average atomic mass of carbon?

C-12 natural abundance = 98.90% , atomic mass = 12 amu

C-13 natural abundance = 1.10% , atomic mass = 13 amu

Note that 98.9 % = $\frac{98.9}{100}$ = 0.989

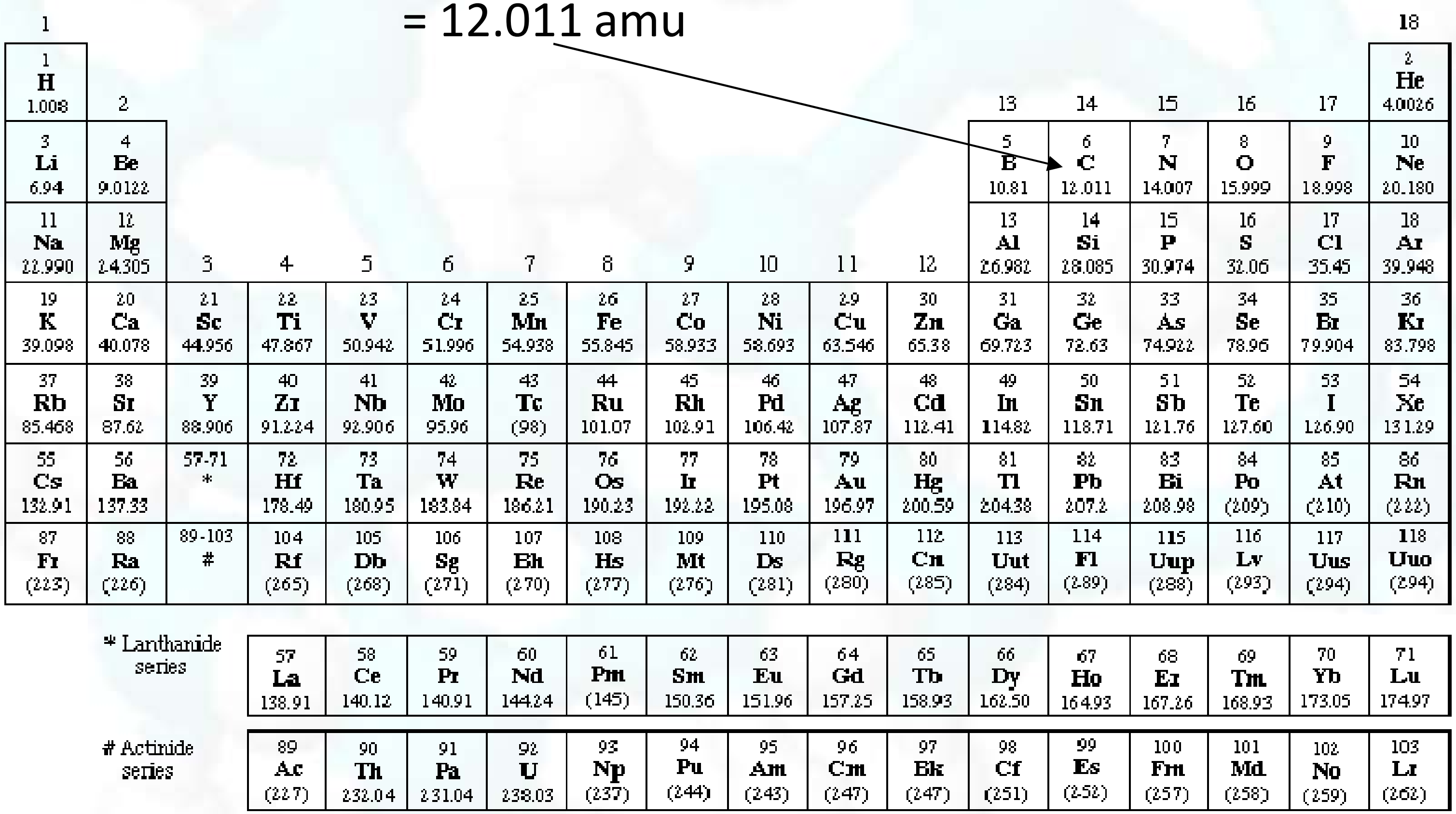
1.10 % = $\frac{1.10}{100}$ = 0.011

Average atomic mass = (natural abundance x atomic mass $)_{C-12}$+ (natural abundance x atomic mass $)_{C-13}$

Average atomic mass

Average atomic mass for C = (0.989 x 12) + (0.011 x 13)

= 12.011 amu



1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.0026
3 **Li** 6.94	4 **Be** 9.0122											5 **B** 10.81	6 **C** 12.011	7 **N** 14.007	8 **O** 15.999	9 **F** 18.998	10 **Ne** 20.180
11 **Na** 22.990	12 **Mg** 24.305											13 **Al** 26.982	14 **Si** 28.085	15 **P** 30.974	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.948
19 **K** 39.098	20 **Ca** 40.078	21 **Sc** 44.956	22 **Ti** 47.867	23 **V** 50.942	24 **Cr** 51.996	25 **Mn** 54.938	26 **Fe** 55.845	27 **Co** 58.933	28 **Ni** 58.693	29 **Cu** 63.546	30 **Zn** 65.38	31 **Ga** 69.723	32 **Ge** 72.63	33 **As** 74.922	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.798
37 **Rb** 85.468	38 **Sr** 87.62	39 **Y** 88.906	40 **Zr** 91.224	41 **Nb** 92.906	42 **Mo** 95.96	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57-71 *	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89-103 #	104 **Rf** (265)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 **Uut** (284)	114 **Fl** (289)	115 **Uup** (288)	116 **Lv** (293)	117 **Uus** (294)	118 **Uuo** (294)

* Lanthanide series	57 **La** 138.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
# Actinide series	89 **Ac** (227)	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Average atomic mass

- Example 2:

Calculate the average atomic mass for Li (^{6}Li (7.42%), ^{7}Li (92.58%))?

Li-6 natural abundance = 7.42% , atomic mass = 6 amu

Li-7 natural abundance = 92.58% , atomic mass = 7 amu

Note that 7.42 % = $\frac{7.42}{100}$ = 0.0742.

92.58 % = $\frac{92.58}{100}$ = 0.9258.

Average atomic mass = (natural abundance x atomic mass $)_{Li-6}$+ (natural abundance x atomic mass $)_{Li-7}$

Average atomic mass for Li = (0.0742 x 6) + (0.9258 x 7)

= 6.94 amu.

Avogadro’s number

- The ***mole (mol)*** is the amount of a substance that contains as many elementary entities as there are atoms in exactly 12.00 grams of C-12.
- They calculate this number and it was found to be $6.022 \times 10^{+23}$.
- This number was called Avogadro's number (N_A) (after Italian scientist Amedeo Avogadro).
- 1 dozen of H_2O = 12 H_2O
- 1 mole of $H_2O = 6.022 \times 10^{+23}$ H_2O atom
- 1 mole of $CO_2 = 6.022 \text{ X } 10^{+23}$ CO_2 atom
- 1 mole of cars = $6.022 \text{ X } 10^{+23}$ of cars
- 1 mole of shoes = $6.022 \text{ X } 10^{+23}$ of shoes

Molar Mass

- Molar mass (M)= the mass (in gram) of 1 mole of a substance.

For any element

atomic mass (amu) = molar mass (grams/mole)

- Example:

Atomic mass of Na is 22.99 amu then Molar mass is 22.99 g/mol

Atomic mass of P is 30.97 amu then Molar mass is 30.97 g/mol

Molar Mass

- To calculate number of particle (atoms or molecules) :

Number of particle = Avogadro's number x number of moles.

Example1:

Calculate the number of atoms in 2 mole of hydrogen?

Number of atoms = 6.022 x 10^{+23} x 2 = 12.044 x 10^{+23}.

Example 2 :

Calculate the number of moles in 6 x 10^{+20} atom of helium (He)?

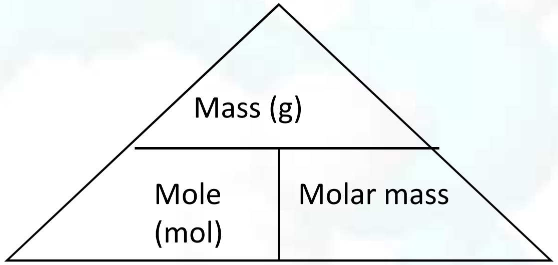
Number of mole = number of atom / avogadro’s number

= 6 x 10^{+20} / 6.022 x 10^{+23}

= 0.99 x 10^{-3} mole.

Molar Mass

- Relationships between mole and molar mass:
- n (number of moles)= mass (g) / molar mass (g/mole)



- Example 1:

How may grams of Zn in 0.356 mole of Zn?

First find the molar mass of Zn from periodic table

65.39 g/mole

Number of mole = mass / molar mass

Mass = number of mole x molar mass

= 0.356 (mole) x 65.39 (g/mole)

= 23.3 g.

Molar Mass

- Example 2 :

Calculate the number of atoms in 6.46 grams of helium (He)?

The molar mass of He is 4 g/mole

First calculate number of moles of He:

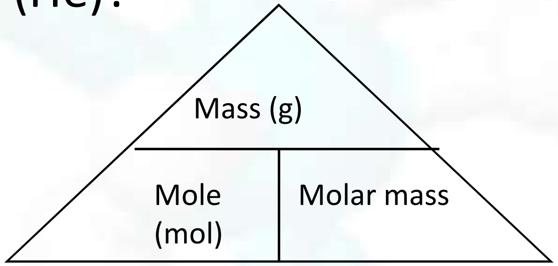
n = 6.46 (g) / 4 (g/mole)

= 1.62 mole

Number of atom = avogadro’s number x number of mole

= $6.022 \times 10^{+23} \times 1.62$

= 9.73×10^{23} atoms



Molecular Mass

- The molecular mass (also called molecular weight) of compound is the sum of all molar mass of each of it elements.

Example1:

what is the Molecular mass for NaOH?

NaOH consist of Na , O, H

Then molecular mass (NaOH)= molar mass of Na + molar mass of O + molar mass of H

= 23 + 16.00 + 1 = 40 amu

Example 2:

What is the Molecular mass of SO_2

Then molecular mass (SO_2)= molar mass of S + 2X(molar mass of O)

= 32 + 2(16)

= 64 amu

Molecular Mass

Example 3:

What is the molecular mass of $C_3H_4O_2$?

Then molecular mass ($C_3H_4O_2$)= 3X(molar mass of C) + 4X(molar mass of H) + 2X(molar mass of O)

= 3 X12 + 4X 1 + 2X 16

= 72 amu

- Molecular mass (amu) = molar mass (g/mole) .

Molecular Mass

Example 4:
How many molecules of ethane (C_2H_6) are present in 0.334 g of C_2H_6?

Molar mass = (2 x 12)+ (6 x 1) = 30 g/mole

First we should calculate the number of mole

Number of mole = mass / molar mass

Number of mole = 0.334 / 30 = 0.011 mole

We know that
Number of molecules = Avogadro's number x number of mole
= $6.022 \times 10^{23} \times 0.011$
= 6.624×10^{21} molecules.

Molecular Mass

- Example 5:

How many hydrogen atoms are present in 25.6 g of urea [$(NH_2)_2CO$]. The molar mass of urea is 60.06 g/mol?

First we calculate the number of mole

Number of mole = mass / molar mass

= 25.6 / 60.06 = 0.426 mole

Number of molecules = avogadro's number x number of mole

= $6.022 \times 10^{23} \times 0.426$

= 2.57×10^{23} molecules

From the chemical formula of urea $(NH_2)_2CO$

1 molecules of urea = 4 atom of H

2.567×10^{23} molecules = ? Atom oh H

$4 \times 2.567 \times 10^{23} = 1.03 \times 10^{24}$ atoms

Molecular Mass

- Example 6:

What is the mass, in grams, of one copper atom?

I know that

1 mole of Cu ====== 6.022 x 10^{23} atom of Cu

Molar mass of Cu= 63.55 g/mol

That mean 63.55 g of Cu ======= 1mole of Cu

Thus

1 mole of Cu ====== 6.022 x 10^{23} atom of Cu

1 mole of Cu ====== 63.55 g of Cu

Then

6.022 x 10 ^{23} atom of Cu ====== 63.55 g of Cu

1 atom of Cu =======? g of Cu → او س

6.22 X10^{23} x ?g= 63.55 x 1

$$\text{grams of Cu} = \frac{1 \times 63.55}{6.022 \times 10^{23}} = \frac{10.55}{10^{23}} = 10.55 \times 10^{-23}\ \text{g}$$

Mass of one atom = molar mass / Avogadro number

Percent Composition of a Compounds

- Percent composition by mass is the percent by mass of each element in a compound.
- This can be obtained from this formula:

$$\frac{n \text{ x molar mass of element}}{\text{molar mass of compound}} \text{ x } 100\%$$

- Where n is the number of moles of the element in 1 mole of compound

Percent Composition of a Compounds

Example 1:

Calculate the percentage of each element in H_2O_2?

First we calculate the molar mass of the compound

(2 x 1) + (2 x 16) = 34 g/mol

$$\% \text{ of H} = \frac{2 \times 1}{34} \times 100\% = 5.88\ \%$$

$$\% \text{ of O} = \frac{2 \times 16}{34} \times 100\% = 94.12\ \%$$

THE SUM SHOULD BE 100

5.88 + 94.12 = 100%

Percent Composition of a Compounds

Example 2:

Phosphoric acid (H_3PO_4) is a colorless , syrupy liquid used in detergents, fertilizers, toothpastes, and in carbonates beverages for a “tangy” flavor. Calculate the percent composition by mass of H, P, and O in this compound?

First we calculate the molar mass of the compound

(3 x 1) + 31 +(4 x 16) =98 g/mol

$$\% \text{ of H} = \frac{3\text{x } 1}{98} \text{ x } 100\% = 3.06\ \%$$

$$\% \text{ of P} = \frac{1\text{x}31}{98} \text{ x } 100\% = 31.63\ \%$$

$$\% \text{ of O} = \frac{4 \text{ x } 16}{98} \text{ x } 100\% = 65.31\ \%$$

THE SUM SHOULD BE 100

3.06 + 65.31 + 31.63 = 100%

Percent Composition of a Compounds

- Example 3:

Which of these is the richest source of nitrogen on a mass percentage basis?

A- Urea, $(NH_2)_2CO$

b- Ammonium nitrate, NH_4NO_3

C- Guanidine, $HNC(NH_2)_2$

d- Ammonia, NH_3^+

We have to calculate the percentage of Nitrogen in each compound then compare between them, the one who have the highest percentage of nitrogen is the richest source of nitrogen.

Percent Composition of a Compounds

- It possible to determine the empirical formula from the percentage of elements in the compound.

1- change % to g

2- change g to mole (remember the triangle).

3- divide by the smallest number of moles.

4- if there was fraction after division change to integer subscripts (multiply by 1 or 2 or 3 etc until reach integer.

Percent Composition of a Compounds

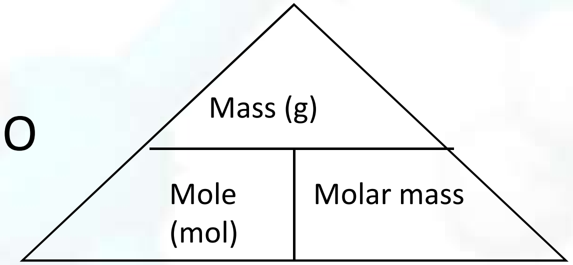
Example 1:

Determine the empirical formula of Vitamin C. it is compose of 40.92% of C, 4.58% of H, and 54.50% of O by mass?

1- change from % to g

40.92 g of C, 4.58 g of H , 54.50 g of O

2- change from g to mole using



$$n_c = \frac{40.92}{12} = 3.41 \text{ mol of C}$$

$$n_H = \frac{4.58}{1} = 4.58 \text{ mol of H}$$

$$n_O = \frac{54.50}{16} = 3.406 \text{ mol of O}$$

Divided by the smallest number of mole which is 3.406

Percent Composition of a Compounds

C: $\frac{3.41}{3.406} \approx 1$ H: $\frac{4.58}{3.406} = 1.34$ O: $\frac{3.406}{3.406} = 1$

4- Because number of hydrogen is 1.33 then we start to multiply until we reach integer, this is trail and error procedure:

1 x 1.34 = 1.34

2 x 1.34 =2.68

3 x 1.34 = 4.02 ≈ 4

Then we multiply all element with the same number we stopped at (which is 3 in this example).

C : 1x3 =3 , O: 1 x 3 = 3 , H =4

Thus the empirical formula is $C_3H_4O_3$

Percent Composition of a Compounds

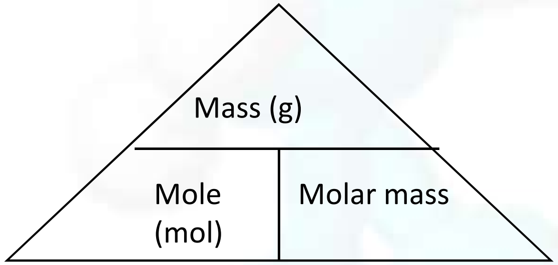
- Example 2 :

Allicin is the compound responsible of characteristic smell of garlic. an analysis of the compound gives the following percent composition by mass: C: 44.4 %, H: 6.21%, S: 39.5%, O: 9.86%. Calculate its empirical formula?

1- change from % to g

44.4 g of C, 6.21 g of H , 39.5 g of S, 9.86 g of O.

2- change from g to mole using


Mass (g)
Mole (mol)
Molar mass

$$n_C = \frac{44.4}{12} = 3.70 \text{ mol of C} \qquad n_H = \frac{6.21}{1} = 6.21 \text{ mol of H}$$

$$n_S = \frac{39.5}{32} = 1.23 \text{ mol of S} \qquad n_O = \frac{9.86}{16} = 0.62 \text{ mol of O}$$

3- Divide by the smallest number of mole which is 0.62

Percent Composition of a Compounds

$$C: \frac{3.70}{0.62} \approx 6 \qquad H: \frac{6.21}{0.62} \approx 10$$

$$S: \frac{1.32}{0.62} \approx 2 \qquad O: \frac{0.62}{0.62} = 1$$

Because all the numbers are integer then we do not need to do anything else and we just write the formula as following :

The empirical formula is $C_6H_{10}O\ S_2$

Percent Composition of a Compounds

- It also possible to determine the molecular formula from the percentage of elements in the compound.

1-determin the empirical formula (as before)

a-change % to g

b- change g to mole (remember the triangle).

c- divide by the smallest number of moles.

d- if there was fraction after division change to integer subscripts (multiply by 1 or 2 or 3 etc until you reach integer.

2- calculate the molecular mass of empirical formula

3- calculate the ratio between molecular formula and empirical formula as following:

$$\text{Ratio} = \frac{\text{molar mass of compound known form the quesition}}{\text{empirical molar mass}}$$

4-Then multiply the empirical formula with that ratio number

Percent Composition of a Compounds

Example 1:

1-Determine the molecular formula of Vitamin C. it is compose of 40.92% of C, 4.58% of H, and 54.50% of O by mass and the molar mass of vitamin C is 176 g/mol?

First we determine the Empirical formula as seen before for this compound

Empirical formula = $C_3H_4O_3$

2- calculate the molar mass of empirical formula:

3 x 12 + 4 x 1 + 3 x16 = 88 g/mol

3- calculate the ratio

$$\text{Ratio} = \frac{\text{molar mass of compound}}{\text{empirical molar mass}}$$

$$\text{Ratio} = \frac{176}{88} = 2$$

4- molecular formula = ratio x empirical formula

= 2 x $C_3H_4O_3$ = $C_6H_8O_6$

Percent Composition of a Compounds

Example 2:

Allicin is the compound responsible of characteristic smell of garlic. an analysis of the compound gives the following percent composition by mass: C: 44.4 %, H: 6.21%, S: 39.5%, O: 9.86%. Calculate its molecular formula if you know that the molar mass of the compound is 162.27 g/mol?

First we determine the empirical formula as seen before for this compound

Empirical formula = $C_6H_{10}O\ S_2$

2- calculate the molar mass of empirical formula:

6 x 12 + 10 x 1 + 16 + 2 X 32 = 162 g/mol

3- calculate the ratio

$$\text{Ratio} = \frac{\text{molar mass of compound}}{\text{empirical molar mass}}$$

$$\text{Ratio} = \frac{162.27}{162} = 1$$

4- molecular formula = ratio x empirical formula

= 1 x $C_6H_{10}O\ S_2$ = $C_6H_{10}O\ S_2$

Chemical Reactions and Chemical Equations

- Often chemist also indicate the physical state of the reactants and products by using the letters *g, l, s.* and when chemical is dissolved in water the symbol (*aq*) is used.

$$O_{2(g)} + 2H_{2(g)} \longrightarrow H_2O_{(l)}$$

$$KBr_{(aq)} + AgNO_{3(aq)} \longrightarrow KNO_{3(aq)} + AgBr_{(s)}$$

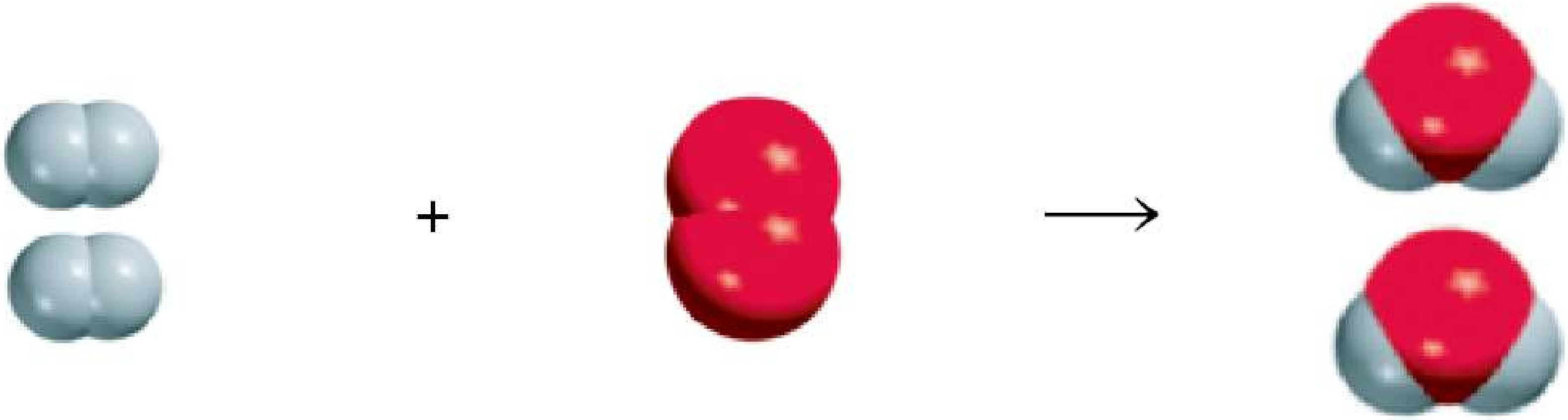
Chemical Reactions and Chemical Equations

- Chemical reaction is a process in which a substance (or substances) is changed into one or more new substances.
- Chemical equation uses chemical symbols to show what happens during a chemical reaction.
- Remember in chemical reaction we don't create new elements.
- The chemicals that react with each other called reactant, and the chemicals produced from the reaction is called product.
- To write a chemical equation we always put the reactants on left side and the products on the wright side

$$\text{Reactants} \longrightarrow \text{products}$$

Chemical Reactions and Chemical Equations

There are 3 way to represent the reaction between hydrogen and oxygen:



Two hydrogen molecules + One oxygen molecule ⟶ Two water molecules

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Chemical Reactions and Chemical Equations

- How to read chemical equations

$$2Mg + O_2 \longrightarrow 2MgO$$

a) 2 atom of Mg + 1 molecules of O_2 makes 2 molecules of MgO

b) 2 mole of Mg + 1 mole of O_2 makes 2 mole of MgO

C) From perodic table

48.6 g of Mg + 32 g of O_2 makes 80.6 g of MgO

Molar mass (g/mol)

d) 2 g of Mg + 1 g of O_2 makes 2 g of MgO ✗

Balancing Chemical Equation

1. Identify all reactants and products and write their correct formula on the left side and right side of the equation.
2. Begin balancing by Change the numbers in front of the formulas (***coefficients***) to make the number of atoms in each element the same on both sides of the equation. Do not change the subscripts.

   NO_2 when multiply by 2=======$2NO_2$ not N_2O_4
3. Start by balancing those elements that appear in only one reactant and one product.
4. Balance those elements that appear in two or more reactants or products.
5. Check to make sure that you have the same number of each type of atom on both sides of the equation.

Chemical Reactions and Chemical Equations

- Example1 :

Balance the following equation:

$$C_5H_{12} + O_2 \longrightarrow CO_2 + H_2O$$

1. Identify all reactants and products and write their correct formula on the left side and right side of the equation.

Chemical Reactions and Chemical Equations

2. Start by balancing those elements that appear in only one reactant and one product.

So we should start with C or H but not with O

$$C_5H_{12} + O_2 \longrightarrow CO_2 + H_2O$$

5 carbon
on left

1 carbon
on right

Multiply CO_2
by 5

$$C_5H_{12} + O_2 \longrightarrow 5CO_2 + H_2O$$

12 hydrogen
on left

2 hydrogen
on right

Multiply H_2O
by 6

$$C_5H_{12} + O_2 \longrightarrow 5CO_2 + 6H_2O$$

Chemical Reactions and Chemical Equations

3. Balance those elements that appear in two or more reactants or products

$$C_5H_{12} + O_2 \longrightarrow 5CO_2 + 6H_2O$$

Multiply O_2 by 8

2 oxygen on left

5x2 oxygen on right

6 oxygen on right

= 16 oxygen on right

$$C_5H_{12} + 8O_2 \longrightarrow 5CO_2 + 6H_2O$$

Chemical Reactions and Chemical Equations

4. Check to make sure that you have the same number of each type of atom on both sides of the equation.

$$C_5H_{12} + 8O_2 \longrightarrow 5CO_2 + 6H_2O$$

C 5	C 5X1 =5
H 12	H 6X2 =12
O 8X2=16	O 5X2 + 6 = 16

Reactants	Products
5 C	5 C
12 H	12 H
16 O	16 O

Chemical Reactions and Chemical Equations

- Example2 :

Balance the following equation:

$$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$$

1. Identify all reactants and products and write their correct formula on the left side and right side of the equation.

Chemical Reactions and Chemical Equations

2. Start by balancing those elements that appear in only one reactant and one product.

So we should start with C or H but not with O

$$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$$

2 carbon
on left

1 carbon
on right

Multiply CO_2
by 2

$$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$$

6 hydrogen
on left

2 hydrogen
on right

Multiply H_2O
by 3

$$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$$

Chemical Reactions and Chemical Equations

3. Balance those elements that appear in two or more reactants or products

$$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$$

Multiply O_2 by 7/2

2 oxygen on left

2x2 oxygen on right

3 oxygen on right

= 7 oxygen on right

$$C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$$

remove fraction by multiply both sides by 2

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

Chemical Reactions and Chemical Equations

4. Check to make sure that you have the same number of each type of atom on both sides of the equation.

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

C 2X2=4	C 4X1 =4
H 2X6=12	H 6X2 =12
O 7X2=14	O 4X2 + 6 = 14

Reactants	Products
4 C	4 C
12 H	12 H
14 O	14 O

Chemical Reactions and Chemical Equations

- Example3 :

Balance the following equation:

$$Al + O_2 \longrightarrow Al_2O_3$$

1. Identify all reactants and products and write their correct formula on the left side and right side of the equation.

Chemical Reactions and Chemical Equations

2. Start by balancing those elements that appear in only one reactant and one product.

All two element (Al,O) appear only once on each side so we can start with any one.

$Al + O_2 \longrightarrow Al_2O_3$

Multiply Al by 2

1 aluminum on left

2 aluminum on right

$2Al + O_2 \longrightarrow Al_2O_3$

Multiply O_2 by 3/2

2 oxygen on left

3 oxygen on right

$$2Al + \frac{3}{2}O_2 \longrightarrow Al_2O_3$$

Chemical Reactions and Chemical Equations

$$2Al + \frac{3}{2} O_2 \longrightarrow Al_2O_3$$

remove fraction by multiply both sides by 2

$$4Al + 3O_2 \longrightarrow 2Al_2O_3$$

4. Check to make sure that you have the same number of each type of atom on both sides of the equation.

$$4Al + 3O_2 \longrightarrow 2Al_2O_3$$

Al 4	Al4
O 3X2=6	O 2x3= 6

Reactants	Products
4 Al	4 Al
6 O	6 O

Amount of Reactants and Products

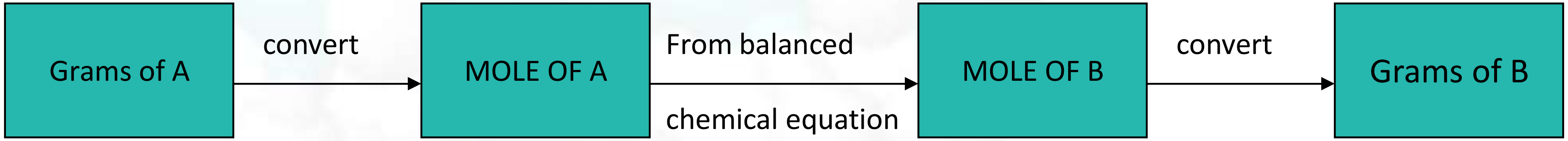
- A basic question in chemical laboratory is How much product will be formed from specific amounts of starting materials (reactant)? Or how much starting materials must be used to obtain a specific amount of product?
- To do that you have to follow the following rules.

1- write the **balanced** equation for the reaction

2- convert the given amount of reactant to moles

3- use the mole ratio from the balanced equation to calculate the number of moles of product.

4- convert the number of moles of product to grams.



Grams of A → convert → MOLE OF A → From balanced chemical equation → MOLE OF B → convert → Grams of B

Amount of Reactants and Products

Example 1:

The food we eat is degraded, or broken down, in our bodies to provide energy for growth and function. A general overall equation for this very complex process represents the degradation of glucose ($C_6H_{12}O_6$) to carbon dioxide and water:

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$$

If 856 g of $C_6H_{12}O_6$ is consumed by a person over a certain period, what is the mass of CO_2 produced?

First we make sure that the equation is balanced

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$$

C 6	C 6
H 12	H 6X2=12
O 6 + 6X2=18	O 6X2 + 6 =18

Amount of Reactants and Products

2- convert g to mole of glucose

$$n = \frac{mass(g)}{molar\ \text{mass (g/mol)}}$$

Then we have to calculate the molar mass of glucose from periodic table

6x12 + 12x1 + 6x 16 = 180 g/mol

$$n = \frac{856\ g}{180\ (g/mol)} = 4.76\ \text{mol}\ C_6H_{12}O_6$$

3- from the equation

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$$

1 mole $C_6H_{12}O_6$ ========= 6 mole of CO_2

4.76 mole $C_6H_{12}O_6$ ========= ? mole of CO_2

1 x ? = 4.76 x 6

$$MOLES\ \text{OF}\ CO_2 = \frac{4.76 \text{ x } 6}{1} = 28.56\ \text{mol of}\ CO_2$$

Amount of Reactants and Products

4- convert mole to g

$$mass(\mathrm{g}) = \text{mole x molar mass}$$

Then we have to calculate the molar mass of CO_2 from periodic table
1x12 + 2x 16 = 44 g/mol

$$mass = 28.56 \text{ x } 44 = 1.256 \text{ x } 10^3 \text{ g of } CO_2$$

Amount of Reactants and Products

Example 2 :

All alkali metals reacts with water to produce hydrogen gas and the corresponding alkali metal hydroxide. A typical reaction is that between lithium and water:

$$2Li + 2H_2O \longrightarrow 2LiOH + H_2$$

How many grams of Li are needed to produced 9.89 g of H_2?

First we make sure that the equation is balanced

$$2Li + 2H_2O \longrightarrow 2LiOH + H_2$$

Li 2	Li 2
H 4	H 2+2=4
O 2	O 2

Amount of Reactants and Products

2- convert g to mole of glucose

$$n = \frac{mass(g)}{molar\ \text{mass (g/mol)}}$$

Then we have to calculate the molar mass of hydrogen from periodic table

2 X1 = 2 g/mol

$$n = \frac{9.89}{2} = 4.94\ mol\ \text{of}\ \mathrm{H_2}$$

3- from the equation

$$2Li + 2H_2O \longrightarrow 2LiOH + H_2$$

2 mole Li ========= 1 mole H_2

? mole Li ========= 4.94 mole of H_2

2 X 4.94 = ? X 1

$$mole\ \text{of Li} = \frac{2 \text{ x } 4.94}{1} = 9.88 \text{ mol of Li}$$

Amount of Reactants and Products

4- convert mole to g

$$mass = \text{mole x molar mass}$$

Then we have to calculate the molar mass of Li from periodic table which is 6.941 g/mole

Mass of Li = 9.88 x 6.941 = 68.6 g of Li

Limiting Reagents

$$A + B \longrightarrow C + D$$

- What is limiting reagents

Limiting reagents is the reactant used up first in a reaction.

Excess reagents is the reactant present in quantities greater than necessary to react with the quantity of the limiting reagent (the one that is left at the end of the reaction).

By knowing the limiting reagent we can determine the amount of product.

Always take the smallest number

Limiting Reagents

Example :

Urea is [$(NH_2)_2CO$] is prepared by reacting ammonia with carbon dioxide:

$$2NH_3 + CO_2 \longrightarrow (NH_2)_2CO + H_2O$$

In one process, 637.2g of NH_3 are treated with 1142g of CO_2.

a) Which of the two reactants is the limiting reagent ?

b) Calculate the mass of $(NH_2)_2CO$ formed?

c) How much excess reagent (in grams) is left at the end of the reaction?

a) We should calculate how much product each reactant produced and the one with the smallest number is the limiting reagents.

to do so, we have to use the previous method of calculation.

Limiting Reagents

$$2NH_3 + CO_2 \longrightarrow (NH_2)_2CO + H_2O$$

- First start with NH_3

1-Convert g to mole :

n = 637.2 / 17 = 37.48 mol

2- from equation

2 mole NH_3 ========= 1 mole $(NH_2)_2CO$

37.48 mole NH_3 =====? Mole $(NH_2)_2CO$

1 x 37.48 = 2 x ?

Mole of $(NH_2)_2CO$ = 37.48 / 2

=18.74 mole of $(NH_2)_2CO$

- second start with CO_2

1-Convert g to mole :

n = 1142 / 44 = 25.95 mol

2- from equation

1 mole CO_2 ========= 1 mole $(NH_2)_2CO$

25.95 mole CO_2 =====? Mole $(NH_2)_2CO$

1 x 25.95 = 1 x ?

Mole of $(NH_2)_2CO$ = 25.95 / 1

= 25.95 mole of $(NH_2)_2CO$

Thus the limiting reagent is NH_3 because it produced the least amount of product

Always take the smallest number

Limiting Reagents

b) Calculate the mass of $(NH_2)_2CO$ formed?

Always take the smallest number

We take the number of mole of product formed from the limiting reagent and then converted to grams

The limiting reagent here is NH_3 and it produce 18.74 mole of $(NH_2)_2CO$

Molar mass of $(NH_2)_2CO$ = 2 x 14 + 4x 1 + 12 + 16 = 60 g/mol

Mass = mole x molar mass

= 18.74 x 60 = 1124.4 g of $(NH_2)_2CO$.

Limiting Reagents

c) How much excess reagent (in grams) is left at the end of the reaction?

We need to know how much excess reagent (in this case CO_2) remain after the reaction.

Excess reagent = initial amount of CO_2 - Reacted amount of CO_2

The initial amount of CO_2 we know it from the question (1142g = 25.95 mole)

We need to know how much reacted and we can do this by comparing CO_2 with product $(NH_2)_2CO$

Always take the smallest number

We know that there are 18.74 mole of $(NH_2)_2CO$

From equation

1 mole CO_2 ======== 1 mole $(NH_2)_2CO$.

? Mole of CO_2 ========== 18.74 mole of $(NH_2)_2CO$.

Then the number of mole of reacted CO_2 is 18.74 mole

Then excess reagents = 25.95 – 18.74 = 7.21 mole

= 7.21 X 44 = 317.24 g of CO_2

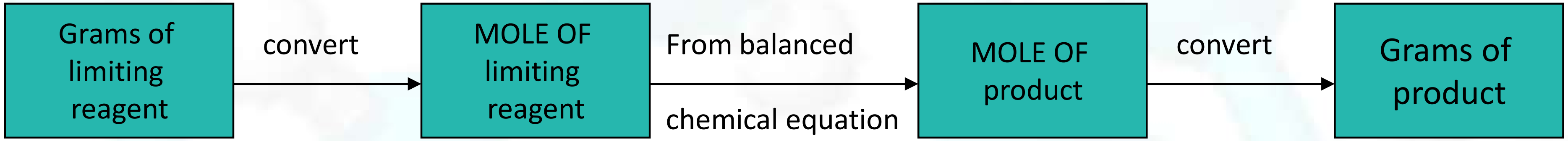
Yield

- For any reaction there are theoretical yield and actual (practical) yield.
- Theoretical yield : the amount of product that would result if all the limiting reagent reacted.
- Actual yield : the amount of product actually obtained from a reaction.
- Normally the actual yield is less than theoretical yield.
- To determine how efficient a given reaction is, we calculate the percent yield.

$$\%\,\text{Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}}\ \text{X}\,100$$

- Normally actual yield is given in the question
- We calculate the theoretical yield from the limiting reagent.

Yield



Grams of limiting reagent → convert → MOLE OF limiting reagent → From balanced chemical equation → MOLE OF product → convert → Grams of product

Example 1:

Titanium is a strong, lightweight, corrosion-resistance metal that is used in rockets, aircraft, jet engines, and bicycle frames. Its prepared by the reaction of titanium (IV) chloride with molten magnesium between 950 °C and 1150 °C:

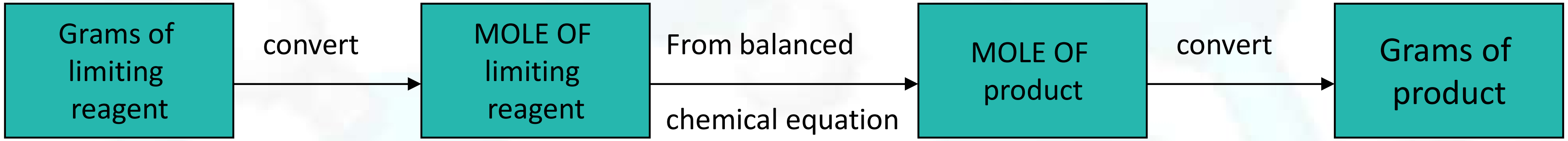
$$TiCl_4 + 2Mg \longrightarrow Ti + 2MgCl_2$$

In a certain industrial operation 3.54 x 10^7 g of $TiCl_4$ are reacted with 1.13 x 10^7 g of Mg.

a) Calculate the theoretical yield of Ti in grams.

b) Calculate the percent yield if 7.91 x 10^6 g of Ti are actually obtained.

Yield



Grams of limiting reagent → convert → MOLE OF limiting reagent → From balanced chemical equation → MOLE OF product → convert → Grams of product

a) Calculate the theoretical yield of Ti in grams.

$TiCl_4 + 2Mg \longrightarrow Ti + 2MgCl_2$

First we have to determine the limiting reagent:

First start with $TiCl_4$:

1-Convert g to mole :

$n = 3.54 \times 10^7 / 189.7 = 1.87 \times 10^5$ mol

2- from equation

1 mole $TiCl_4$ ========= 1 mole Ti

1.87×10^5 mole $TiCl_4$ ========? Mole Ti

$1 \times 1.87 \times 10^5 = 1 \times ?$

Mole of Ti = 1.87×10^5 mole

second start with Mg

1-Convert g to mole :

$n = 1.13 \times 10^7 / 24.31 = 4.65 \times 10^5$ mol

2- from equation

2 mole Mg ========= 1 mole Ti

4.65×10^5 mole Mg ========? Mole Ti

$1 \times 4.65 \times 10^5 = 2 \times ?$

$$\text{mole of Ti} = \frac{1 \text{x} 4.65 \text{ x } 10^5}{2} = 2.32 \text{ x } 10^5 \text{ mol}$$

Mole of Ti = 2.32×10^5 mole

Yield

Thus the limiting reagent is $TiCl_4$ because it produced the least amount of product.

Now we take the number of mole of Ti that produced by the limiting reagent (which is in this example $TiCl_4$)

1.87×10^5 mole of Ti

Convert the mole of TI to g which will be the theoretical yield of Ti

grams of Ti = mole of Ti x molar mass of Ti

$= 1.87 \times 10^5 \times 47.88 = 8.95 \times 10^6$ g of Ti.

Always
Theoretical yield :
mole of limiting agent then convert
to g

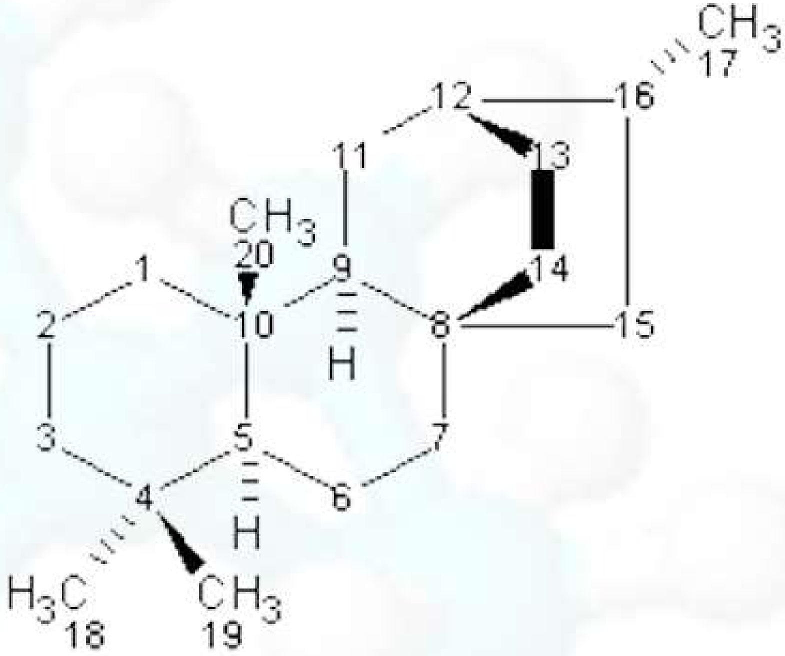
Yield

b) Calculate the percent yield if 7.91 x 10^6 g of Ti are actually obtained.

$$\% \text{ Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \text{ X } 100$$

$$\% \text{ Yield} = \frac{7.91 \text{ x } 10^6}{8.95 \text{ x } 10^6} \text{ X } 100 = 88.4\%$$

Thank you



Chapter Four

Reactions in Aqueous Solutions

Solutions and concentrations

- Solution is a homogenous mixture of two or more substances.
- When water is the solvent, we called the solution aqueous solution.
- Concentration of a solution is the amount of solute present in a given amount of solvent.
- The concentration of a solution can be expressed in many different ways.
- MOLARITY (M): is the number of moles of solute per liter of solution.

$$Molarity = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$M = \frac{\mathrm{n}}{\mathrm{V}}$$

Unit of molarity is mol/L

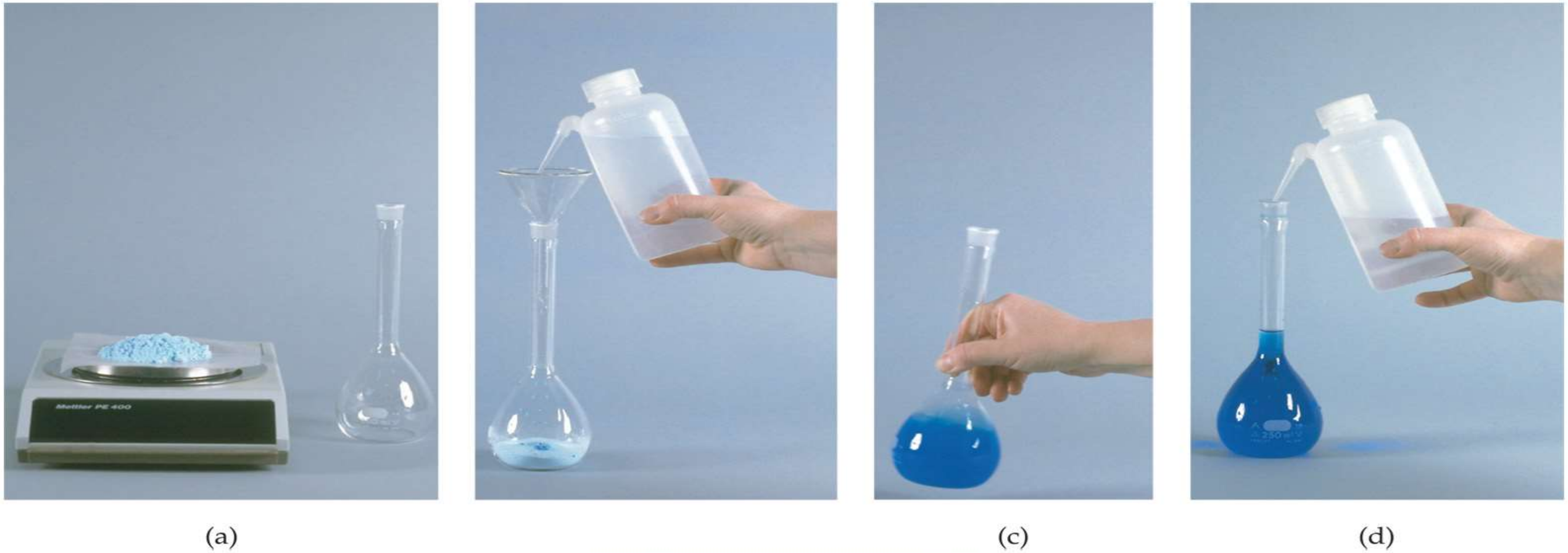
$$where\ \mathrm{M} = \text{molarity}$$

$$\mathrm{n} = \text{number of moles}$$

$$\mathrm{V} = \text{volume of solution in liters}$$

Solutions and concentrations

- Steps to prepare a solution of known molarity :



Solutions and concentrations

Example1:

How many grams of potassium dichromate ($K_2Cr_2O_7$) are required to prepare a 250 ml solution whose concentration is 2.16 M.?

$$M = \frac{n}{V}$$

where M = molarity

n = number of moles

V = volume of solution in liters

M =2.16 , V =250 mL = 250/1000= 0.25 L

n = M X V

= 2.16 X 0.25 = 0.54 mol

n = mass / molar mass

Molar mass of $K_2Cr_2O_7$ = 294.2 g/mol

Mass = n x molar mass

= 0.54 x 294.2 = 158.9 g

Solutions and concentrations

Example 2:

In a biochemical assay, a chemist needs to add 3.81 g of glucose ($C_6H_{12}O_6$) to a reaction mixture. Calculate the volume in milliteres of a 2.53 M glucose solution he should use for the addition.

M = n/V

Molar mass of $C_6H_{12}O_6$ =180 g/mol

n = mass/molar mass

= 3.81 / 180

= 0.021 mol

M = n/V

V = n/M

= 0.021 / 2.53 = 8.30 X 10^{-3} L

= 8.36 X10^{-3} X 10^{3}L = 8.30 mL

Solutions and concentrations

FOR IONIC COMPOUNDS

$NaCl \xrightarrow{\hspace{2cm}} Na + Cl$

1 mol NaCl , 1 mole Na^+ , 1mole Cl^{-1}

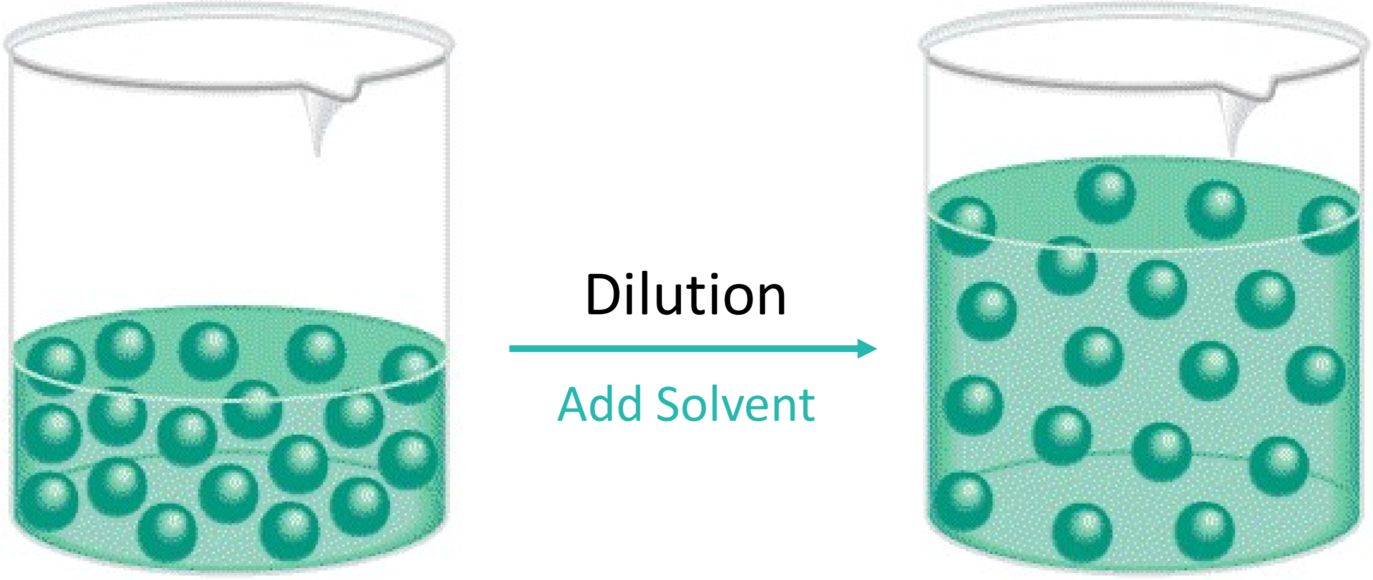
1 M NaCl , 1 M Na^+ , 1 M Cl^{-1}

$Ba\ (NO_3)_2 \xrightarrow{\hspace{2cm}} Ba^{+2} + 2NO_3^-$

1 mole $Ba\ (NO_3)_2$, 1 mole Ba^{+2} , 2 mole NO_3^-

1M $Ba\ (NO_3)_2$, 1 M Ba^{+2}, 2 M NO_3^-

Dilution



Dilution

Dilution : is the procedure for preparing a less concentrated solution from a more concentrated one.

$$M_1 V_1 = M_2 V_2$$

BEFORE AFTER

Example 1 :

How you would prepare 5.00 x 10^2 mL of a 1.75 M H_2SO_4 solution, starting with an 8.61 M stock solution of H_2SO_4?

M_1 = 8.61 , V_1= ?, M_2 = 1.75, V_2 = 5.00 X10^2

$M_1 V_1 = M_2 V_2$

8.61 X V_1 = 1.75 X 5.00 X 10^2

V_1 = 1.75 X 5.00 X 10^2 / 8.61 = 101.6 ml.

Dilution

Example 2:

How would you prepare 60.0 mL of 0.200 *M* HNO_3 from a stock solution of 4.00 *M* HNO_3?

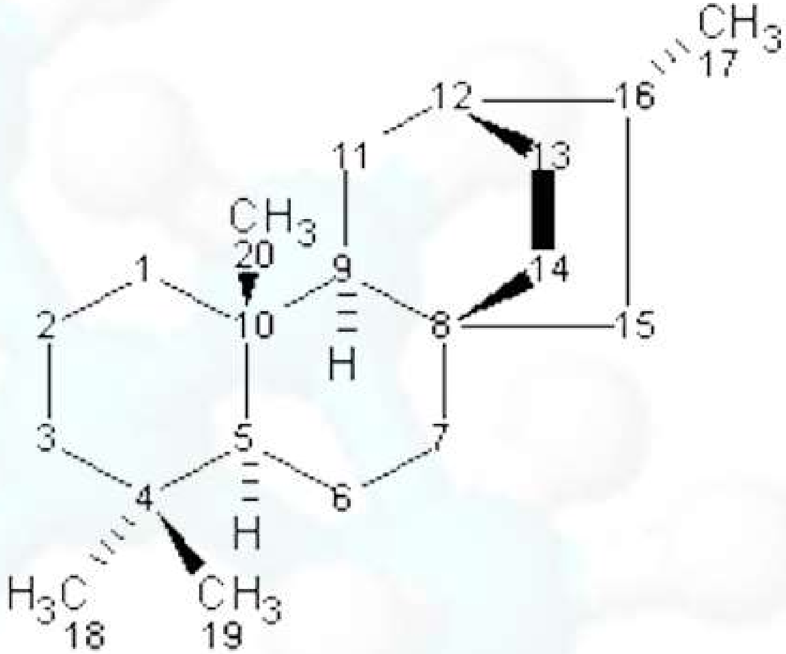
$M_1 = 4$, $V_1 = ?$, $M_2 = 0,2$, $V_2 = 60$

$M_1 V_1 = M_2 V_2$

$4 X V_1 = 0.2 x 60$

$V_1 = 0.2 X 60 / 4 = 3$ ml.

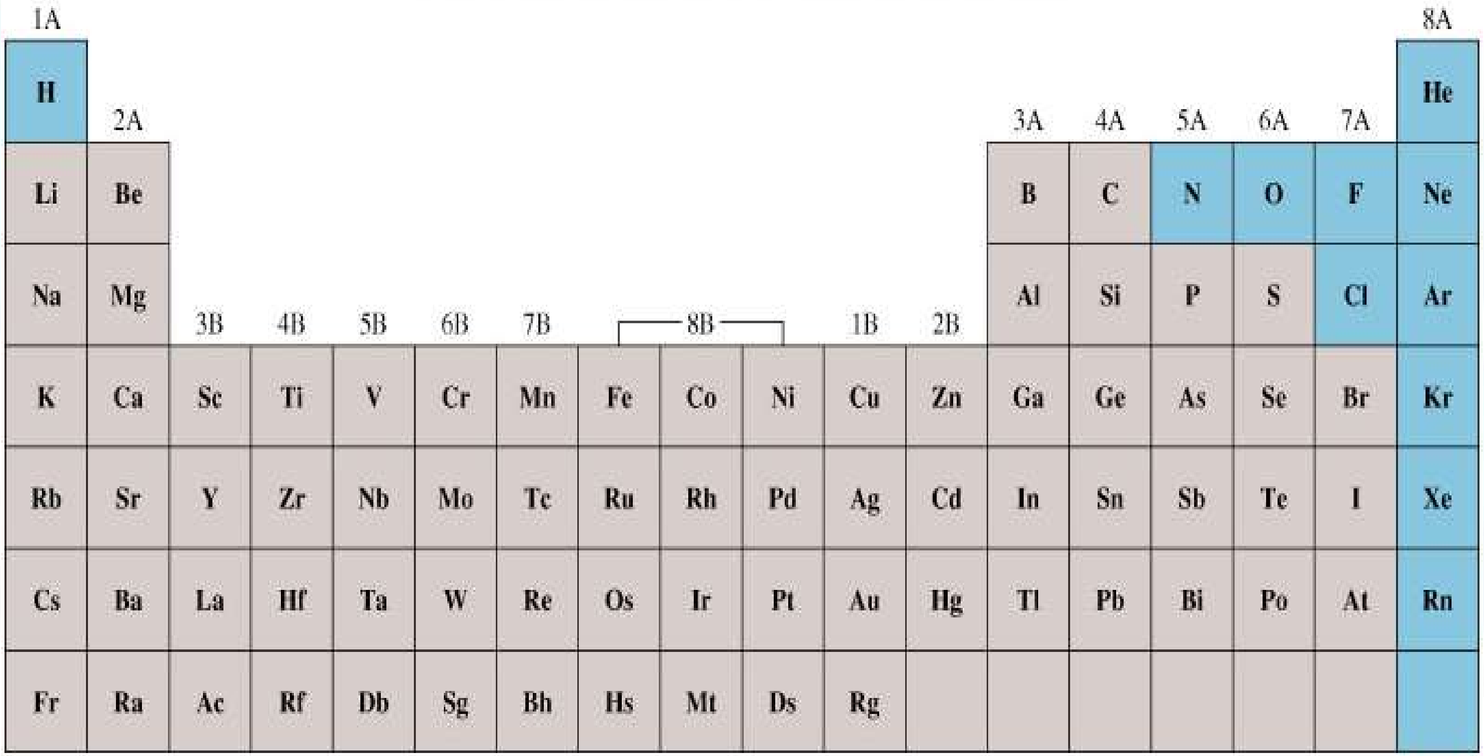
Thank you



Chapter Five

Gases

Substances That Exist as Gases



1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg							

- Element in blue are Gases
- Noble gases are monatomic
- All other gases (H_2, N_2, O_2, F_2, Cl_2) diatomic molecules.

Substances That Exist as Gases

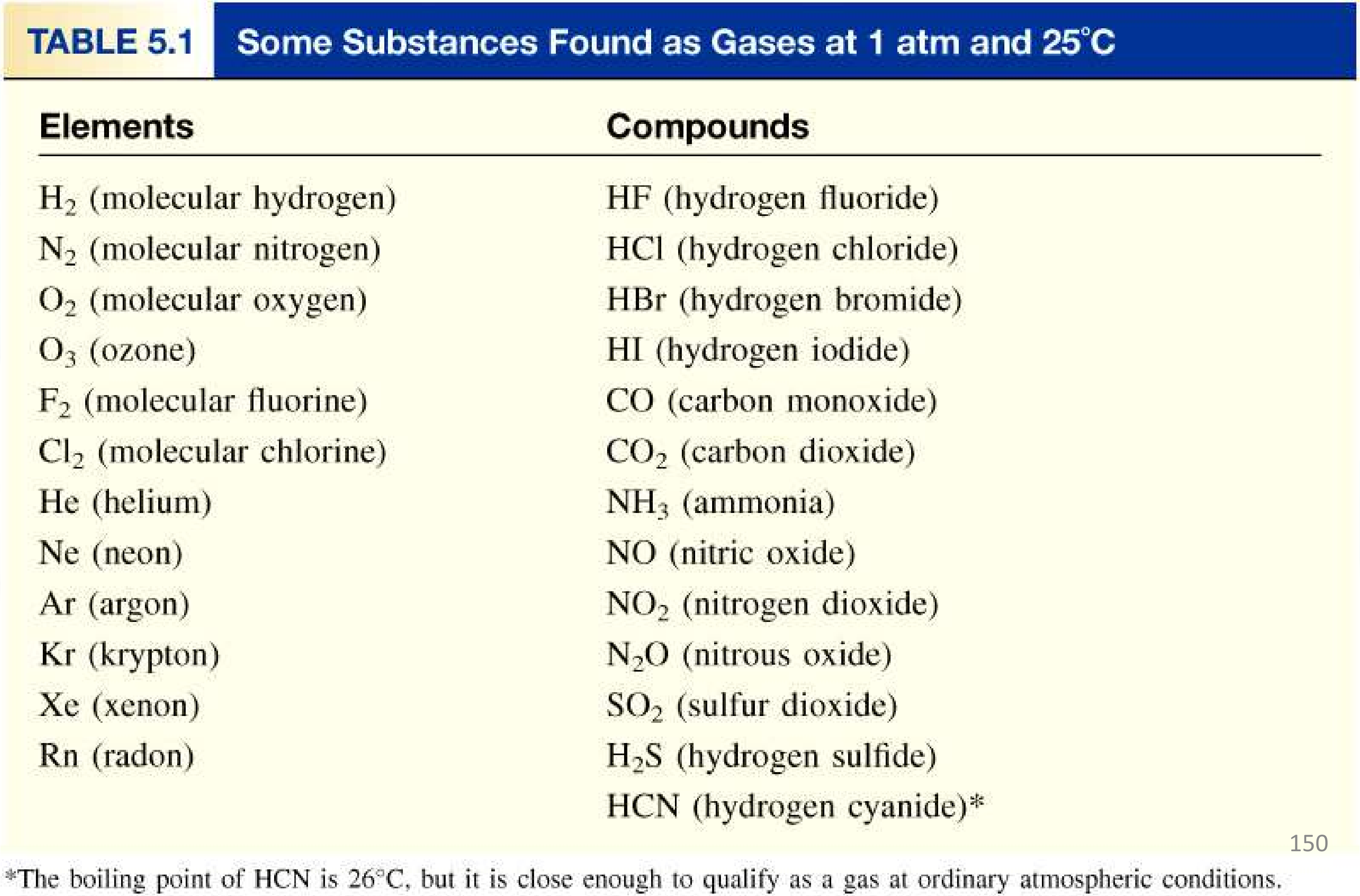


TABLE 5.1 Some Substances Found as Gases at 1 atm and 25°C

Elements	Compounds
H_2 (molecular hydrogen)	HF (hydrogen fluoride)
N_2 (molecular nitrogen)	HCl (hydrogen chloride)
O_2 (molecular oxygen)	HBr (hydrogen bromide)
O_3 (ozone)	HI (hydrogen iodide)
F_2 (molecular fluorine)	CO (carbon monoxide)
Cl_2 (molecular chlorine)	CO_2 (carbon dioxide)
He (helium)	NH_3 (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (argon)	NO_2 (nitrogen dioxide)
Kr (krypton)	N_2O (nitrous oxide)
Xe (xenon)	SO_2 (sulfur dioxide)
Rn (radon)	H_2S (hydrogen sulfide)
	HCN (hydrogen cyanide)*

*The boiling point of HCN is 26°C, but it is close enough to qualify as a gas at ordinary atmospheric conditions.

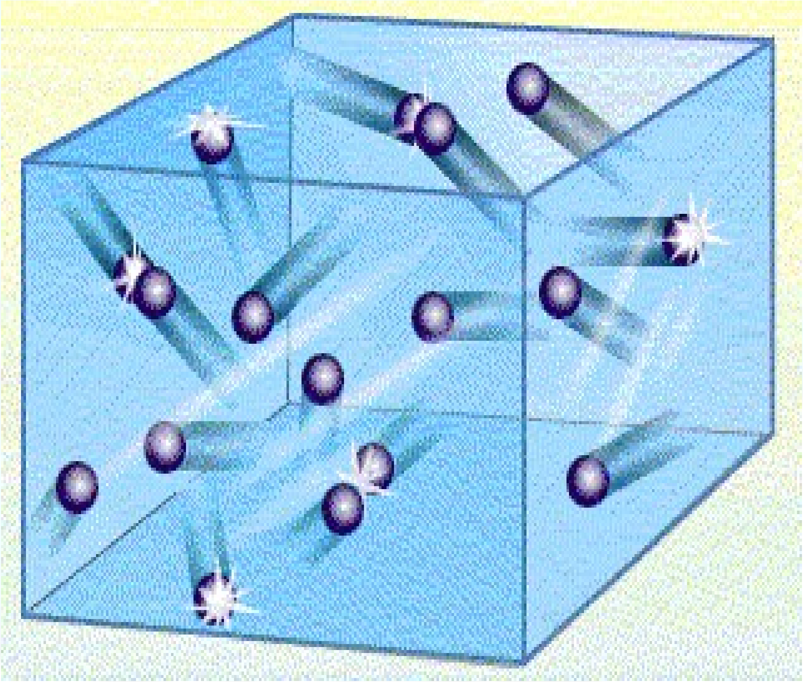
Substances That Exist as Gases

Physical characteristics of Gases :

- Gases assume the volume and shape of their containers.
- Gases are the most compressible of the states of matter.
- Gases will mix evenly and completely when confined to the same container.
- Gases have much lower densities than liquids and solids.

Gas pressure

- Gas particles are in constant moving thus they collide with objects in their bath.
- The gases push against the walls of their containers with a force.
- These collisions produce what we called Gas pressure.

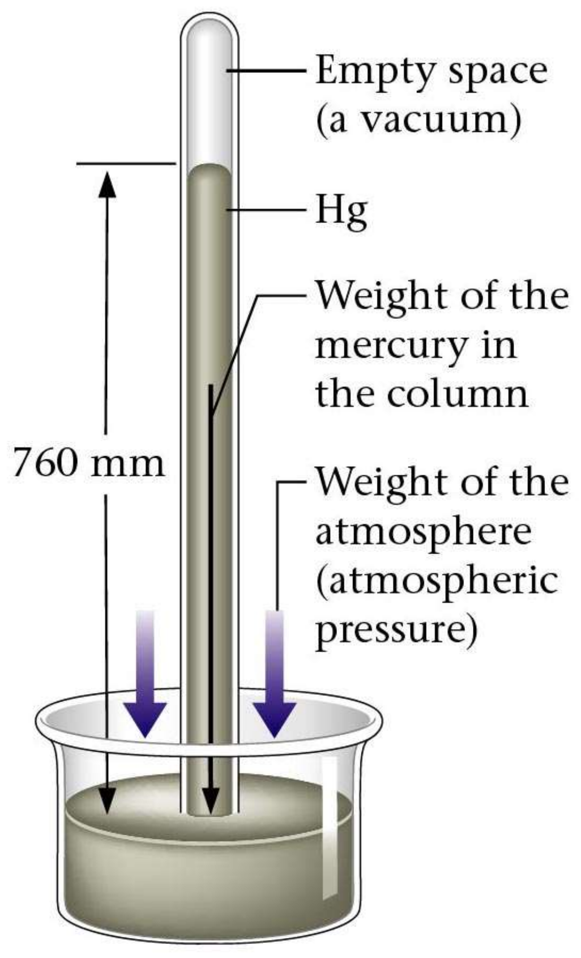


Gas pressure

- Pressure can be defined as :

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

- SI unite of pressure is ***pascal* (Pa)**
- 1 pascal = 1N/m^2. *where N=newton , m=meter*
- Atmospheric pressure is the pressure exerted by Earth's atmosphere.
- We can measure the atmospheric pressure by **barometer**.
- Standard atmospheric pressure (1atm) is the equal to the pressure that supports a column of a mercury exactly 760 mm (or 76cm) high at 0 °C at sea level.
- 1 atm equal a pressure of 760mmHg.



Gas pressure

Unites of pressure :

Pascal (Pa), atm, mmHg, torr

1 torr = 1mmHg

1 atm = 760 mmHg

1 atm = 1.01325×10^5 Pa.

Example1 :

convert the pressure of 688 mmHg to atmospheric pressure?

1 atm = 760mmHg

X

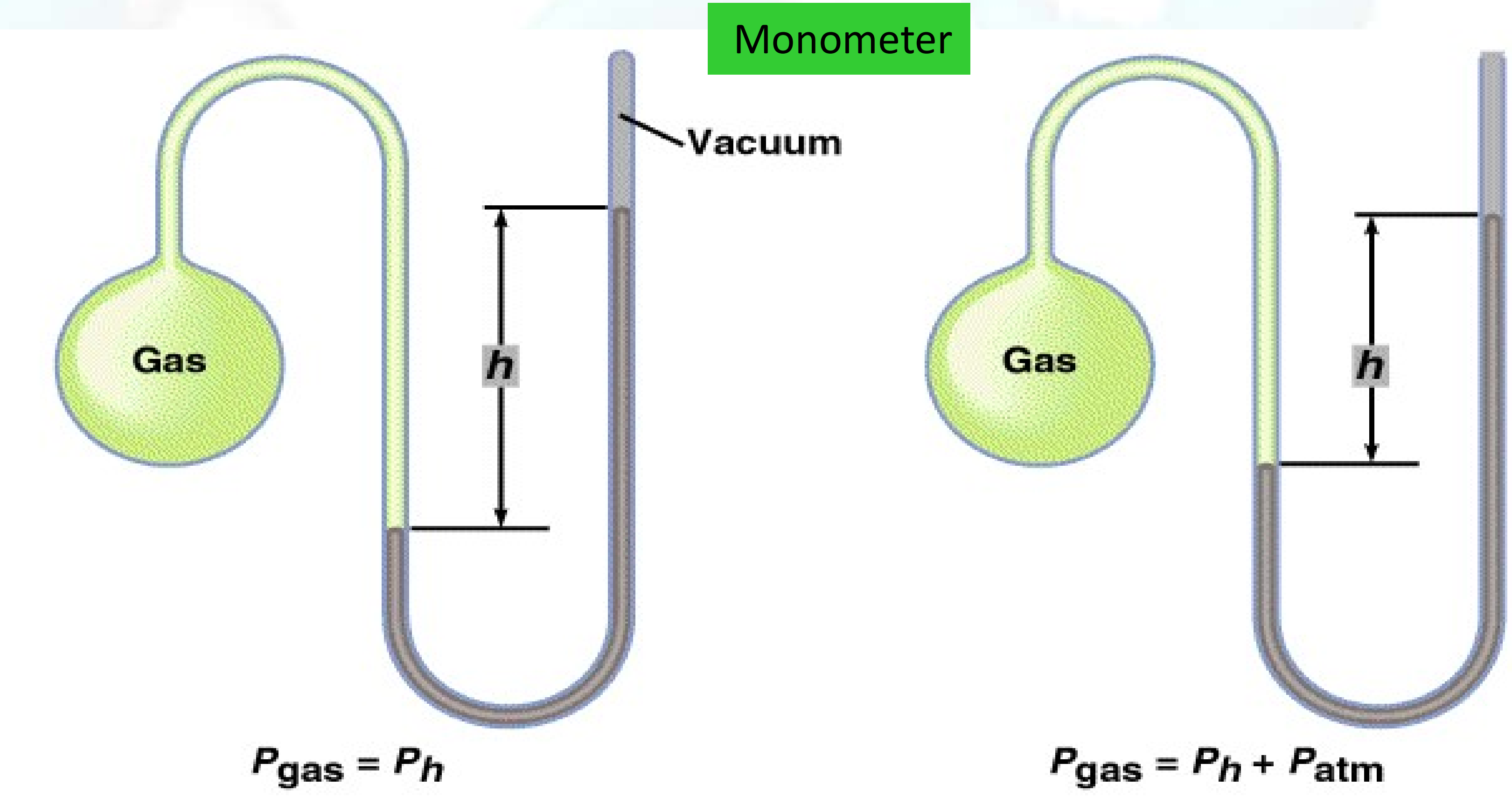
? atm = 688 mmHg

760 X ? = 1 X 688

Pressure = 688 / 760 = 0.905 atm.

Gas pressure

Monometer



$P_{gas} = P_h$

Gas pressure is less than atmospheric pressure

$P_{gas} = P_h + P_{atm}$

Gas pressure is grater than atmospheric pressure

Gas Laws

- For every gas there are :

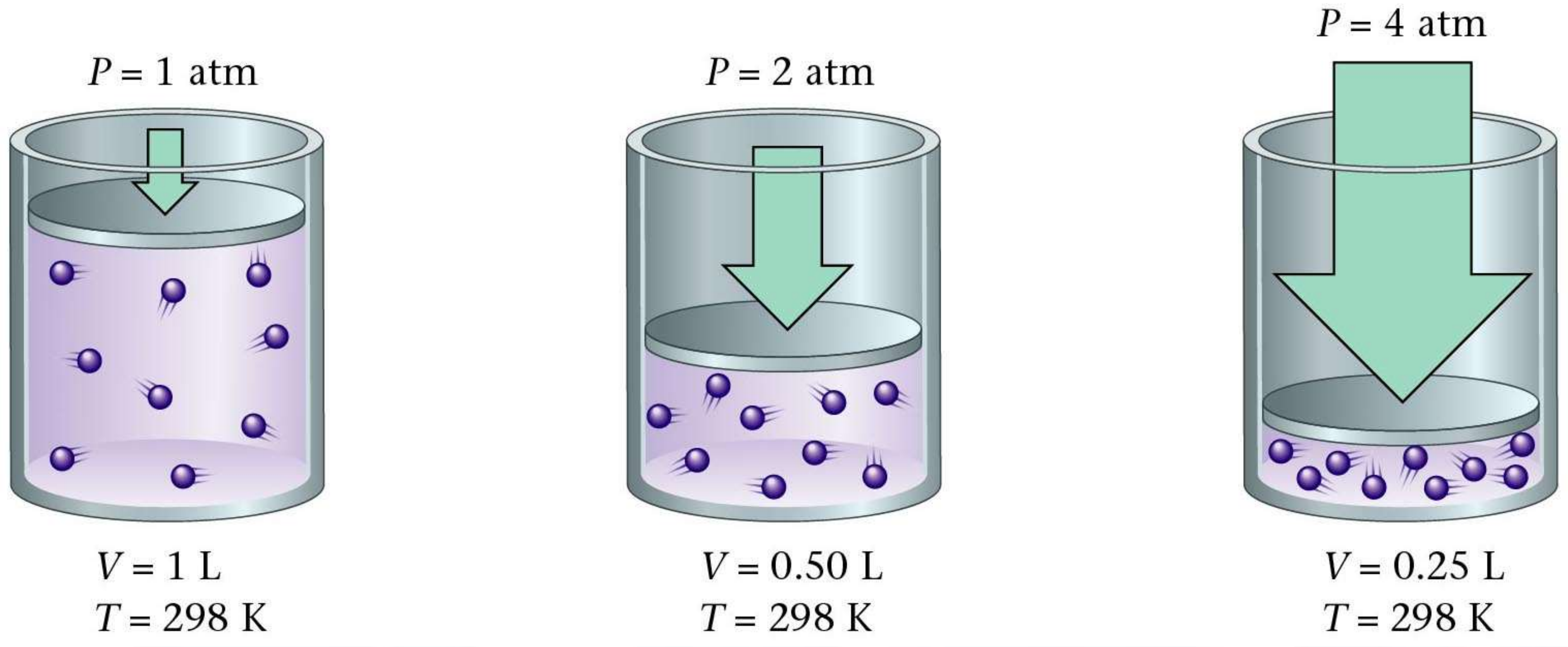
P (pressure), T (Temperature) V (volume), n (mole number).

The Pressure – Volume Relationships
Boyle’s Law

- Boyles’s law study the relationship between the pressure and volume of gas.
- Boyel’s law stated that *the pressure of a fixed amount of gas at a constant temperature in inversely proportional to the volume of the gas.*

$$P \alpha \frac{1}{V}$$

Gas Laws- Boyle's Law



Gas Laws- Boyle's Law

$$P_1 \ \alpha \ \frac{1}{V_1} \quad \Rightarrow P_1 = \frac{k}{V_1} \Rightarrow k = P_1 \text{ x } V_1$$

$$P_2 \ \alpha \ \frac{1}{V_2} \quad \Rightarrow P_2 = \frac{k}{V_2} \Rightarrow k = P_2 \text{ x } V_2$$

$$P_1 V_1 = P_2 V_2$$

Boyle's Law

Gas Laws- Boyle's Law

Example 1 :

A sample of chlorine gas occupies a volume of 946 mL at a pressure of 726 mmHg. What is the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL?

P_1 = 726 mmHg, V_1= 946 ml, P_2= ?, V_2= 154 mL.

$P_1 V_1 = P_2 V_2$

$726 \text{ x } 946 = P_2 \text{ x } 154$

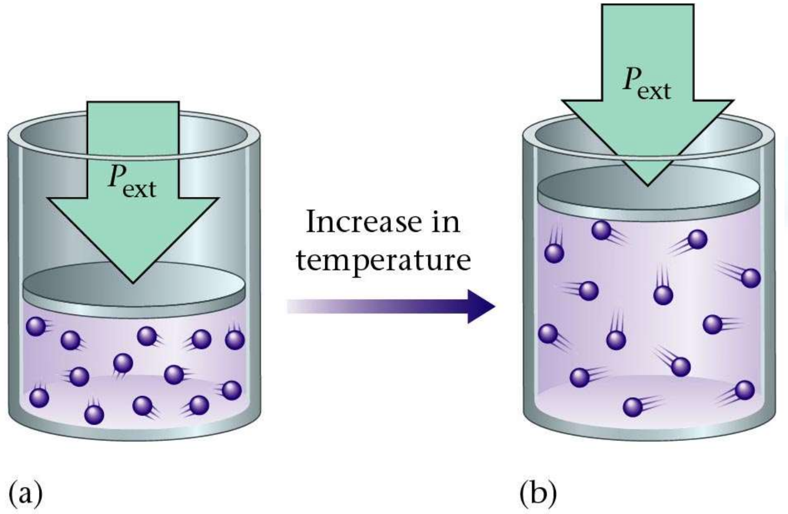
$$P_2 = \frac{726 \text{ x } 946}{154}$$

$$= 4459.7 \text{ mmHg}$$

Gas Laws

The Temperature – Volume Relationships
Charle’s and Gay-Lussac’s Law

- Charle’s and Gay-Lussac’s law study the relationship between the temperature and volume of gas.
- Charle’s and Gay-Lussac’s law stated that the volume of a fixed amount of gas at a constant pressure is directly proportional to the *absolute temperature* of the gas.



Gas Laws- Charle's Law

T α V

$$T_1 \ \alpha \ V_1 \quad \Rightarrow T_1 = k \times V_1 \Rightarrow k = \frac{T_1}{V_1}$$

$$T_2 \ \alpha \ V_2 \quad \Rightarrow T_2 = k \times V_2 \Rightarrow k = \frac{T_2}{V_2}$$

$$\boxed{\frac{T_1}{V_1} = \frac{T_2}{V_2}}$$

Charle's Law

T in Kelvin

Gas Laws- Charle’s Law

Example:

A sample of carbon monoxide gas occupies 3.20 L at 125 °C. At what temperature will the gas occupy a volume of 1.54 L if the pressure remains constant?

$$\frac{T_1}{V_1} = \frac{T_2}{V_2}$$

$$\frac{125 + 273}{3.2} = \frac{T_2}{1.54}$$

T_2 X 3.2 = 398 X 1.54

T_2 = 612.92/3.2

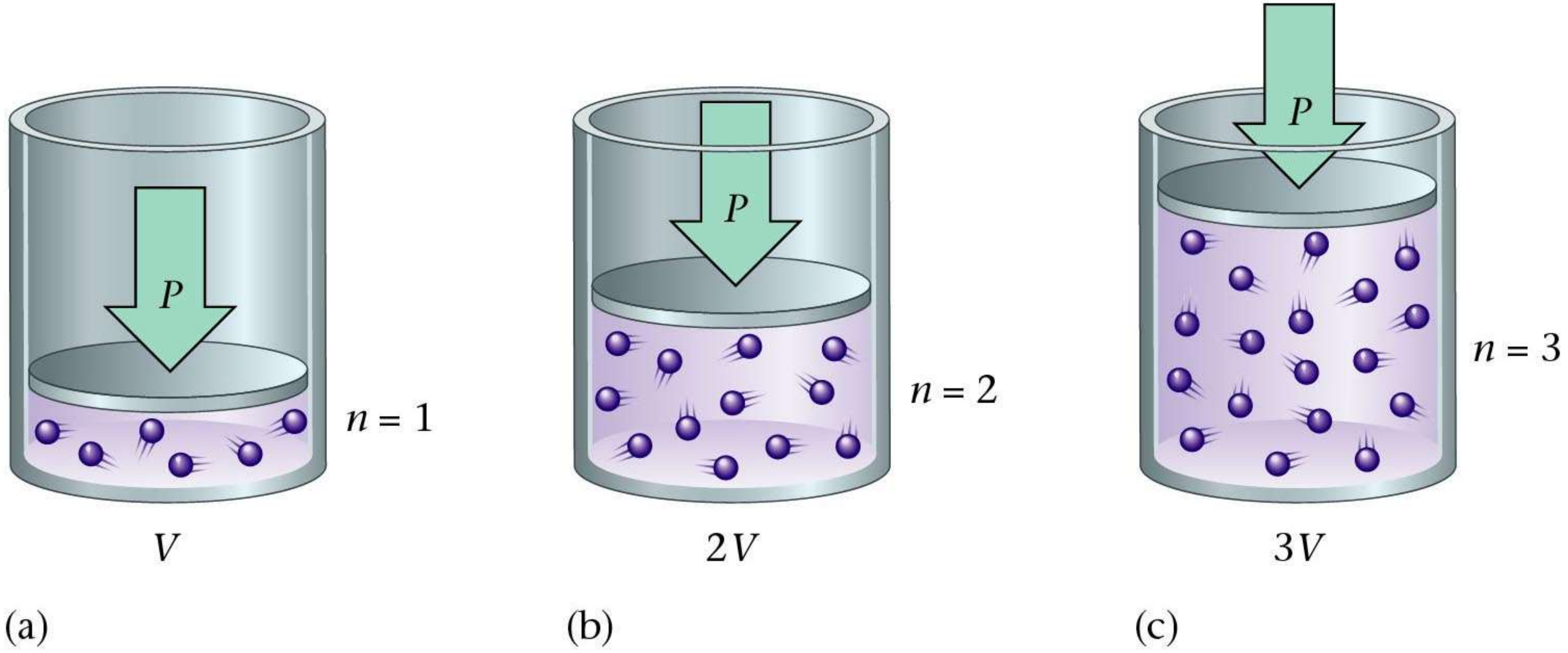
= 191.5 K

Gas Laws

The Volume – Amount Relationships
Avogadro's Law

- Avogadro's law study the relationship between the volume and number of mole of gas.
- Avogadro's law stated that at constant pressure and temperature, the volume is directly proportional to the number of moles of the gas

Gas Laws- Avogadro's Law



Gas Laws- Avogadro's Law

$$n \ \alpha \ V$$

$$n_1 \ \alpha \ V_1 \quad \Rightarrow n_1 = k \ x \ V_1 \Rightarrow k = \frac{n_1}{V_1}$$

$$n_2 \ \alpha \ V_2 \quad \Rightarrow n_2 = k \ x \ V_2 \Rightarrow k = \frac{n_2}{V_2}$$

$$\boxed{\frac{n_1}{V_1} = \frac{n_2}{V_2}}$$

⇐ Avogadro's Law

Summery of Gas Laws

$P_1 V_1 = P_2 V_2$

Boyle's Law
Constant T and n

$$\frac{T_1}{V_1} = \frac{T_2}{V_2}$$

Charle's Law
Constant P and n

$$\frac{n_1}{V_1} = \frac{n_2}{V_2}$$

Avogadro's Law
Constant P and T

Ideal Gas Equation

We know that

$$V \alpha \frac{1}{P} \qquad \text{Boyle's law}$$

$$V \alpha T \qquad \text{Charle's law}$$

$$V \alpha n \qquad \text{A vogadro law}$$

Then

$$V \alpha \frac{nT}{P}$$

$$V = \frac{nRT}{P}$$

Ideal Gas Equation

$$\boxed{PV=nRT}$$

P= pressure (atm), V= volume (L), n= moles
R= gas constant, T= temperature (K)

Ideal Gas Equation

- Ideal gas is a hypothetical gas whose pressure-volume-temperature behavior can be completely accounted for by the ideal gas equation.
- STP : standard Temperature and pressure
- Standard Temperature = 0°C = 273.15 K
- Standard Pressure = 1 atm.
- At STP 1mole of an ideal gas occupies 22.414L.
- R (gas constant) = 0.0821 L.atm / K.mol

Ideal Gas Equation

Example 1:

Calculate the pressure (in atm) exerted by 1.82 moles of the sulphur hexaflouride in a steel vessel of volume 5.43 L at 69.5 °C.?

PV =nRT

P = nRT/V

= 1.82 x 0.0821 x (69.5 +273)/5.43

=9.41 atm.

Ideal Gas Equation

Example 2:

Calculate the volume (in liters) occupied by 7.40g of NH_3 at STP condition.?

PV=nRT

n = mass/molar mass

n = 7.40 / 17 = 0.435 mol

V = nRT/P

V= 0.435 X 0.082 X 273 / 1

= 9.74 L

Ideal Gas Equation

- We can use the ideal gas law if we know three out of four variable namely: P,T,V,n. we can calculate one unknown if we know the other three from the equation of ideal gas.
- However, sometime we have to deal with two conditions, this means we have two P , two V, two T, and two n. thus we need to apply some modification into the equation of ideal gas that take into account the initial and final conditions.

$$PV = nRT \qquad R = \frac{P\,V}{n\,T}$$

$$R = \frac{P_1\,V_1}{n_1\,T_1} \text{ befor change}$$

$$R = \frac{P_2\,V_2}{n_2\,T_2} \text{ after change}$$

$$\frac{P_1\,V_1}{n_1\,T_1} = \frac{P_2\,V_2}{n_2\,T_2}$$

Normally $n_1=n_2$

And the law become

$$\frac{P_1\,V_1}{T_1} = \frac{P_2\,V_2}{T_2}$$

Ideal Gas Equation

Example 1:

A small bubble rises from the bottom of a lake, where the temperature and pressure are 8 °C and 6.4 atm, to the water surface, where the temperature 25 °C and the pressure is 1 atm. Calculate the final volume (in mL) of the bubble if its initial volume was 2.1 mL.

Initial condition	Final condition
P_1 6.4 atm	P_2 1 atm
T_1 8 °C	T_2 25 °C
V_1 2.1 ml	V_2 ?

We assume that air amount in the bubble remains constant ($n_1 = n_2$)

Ideal Gas Equation

Example 1:

A small bubble rises from the bottom of a lake, where the temperature and pressure are 8 °C and 6.4 atm, to the water surface, where the temperature 25 °C and the pressure is 1 atm. Calculate the final volume (in mL) of the bubble if its initial volume was 2.1 mL?

$$\frac{P_1\ V_1}{T_1} = \frac{P_2\ V_2}{T_2}$$

$$P_1 V_1 T_2 = P_2 V_2 T_1$$

$$V_2 = \frac{6.4 \text{ x } 2.1 \text{ x } (25 + 273)}{1 \text{ x } (8 + 273)}$$

V_2 = 14.25 mL

Ideal Gas Equation

Example 2:

An inflated helium balloon with a volume of 0.55L at sea level (1 atm) is allowed to rise to a high of 6.5 km. where the pressure is about 0.40 atm. Assuming that the temperature remains constant. What is the final volume of the balloon?

We assume that $n_1 = n_2$ and $T_1 = T_2$

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

$$P_1 V_1 = P_2 V_2$$

$$1 \text{ x } 0.55 = 0.4 \text{ x } V_2$$

$$V_2 = 0.55 / 0.4$$

$$= 1.4L$$

Density Calculations

$$P\,V = n\,R\,T$$

$$P = \frac{n}{V}R\,T$$

$$\frac{n}{V} = \frac{P}{R\,T}$$

I know that n = mass /molar mass

$$\boxed{\frac{m}{MM\,V}} = \frac{P}{R\,T}$$

I know d= mass/volume

$$\frac{d}{MM} = \frac{P}{R\,T}$$

$$P\,MM = d\,R\,T$$

$$d = \frac{P\,MM}{R\,T}$$

Unite for gas density is g/L

d = density (g/L), P= pressure (atm)
MM= molar mass (g/mol), R= gas constant, T= temperature (K)

Density Calculations

Example 1:

Calculate the density of CO_2 in g/L at 0.990 atm and 55 °C?

$$d = \frac{P\,MM}{R\,T}$$

$MM(CO_2)$= 40 g/mol

$$d = \frac{0.99 \text{ x } 40}{0.0821 \text{ x } (55 + 273)}$$

d= 1.47 g/L

The molar mass of a gaseous substance

- Normally we can determine the molar mass of a compound from the chemical formula.
- However, sometime we work with unknown compound or partially known compound. If the unknown substance is gaseous, its molar mass can be determine from the ideal gas equation. All needed is the density of the gas (or mass and volume of the gas).

$$d = \frac{P\,MM}{R\,T}$$

$$d\,R\,T = P\,MM$$

$$MM = \frac{d\,R\,T}{P}$$

The molar mass of a gaseous substance

Example 1:

A chemist has synthesised a green-yellow gaseous compound of chlorine and oxygen and finds that its density is 7.71 g/L at 36 °C and 2.88 atm. Calculate the molar mass of the compound?

$$\mathrm{MM} = \frac{\mathrm{d\,R\,T}}{\mathrm{P}}$$

$$\mathrm{MM} = \frac{7.71 \text{ x } 0.0821 \text{ x } (36 + 273)}{2.88}$$

$$\mathrm{MM} = 67.9 \text{ g/mol}$$

The molar mass of a gaseous substance

Example 2:

Chemical analysis of a gaseous compound showed that it contained 33.0 percent Si and 67.0 percent F by mass. At 35 °C, 0.210 L of the compound exerted a pressure of 1.70 atm. If the mass of 0.210 L of the compound was 2.38 g, calculate the molar mass and determine the molecular formula of the compound?

Si = 33%, F= 67%, T= 35 °C, V= 0.210L, P= 1.7 atm, mass= 2.38g,

MM= ?, Molecular formula ??

$$MM = \frac{d\,R\,T}{P}$$

$$d = \frac{m}{V}$$

$$d = \frac{2.38}{0.210}$$

d=11.33 g/L

$$MM = \frac{11.33 \text{ x } 0.0821 \text{ x } (35 + 273)}{1.70}$$

$$\boxed{MM=168.5\ g/mol}$$

The molar mass of a gaseous substance

1- change from % to g

33 g of Si, 67 g of F,

2- change from g to mole using

$$n_{Si} = \frac{33}{28.09} = 1.17 \text{ mol of Si}$$

$$n_F = \frac{67}{19} = 3.53 \text{ mol of F}$$

Divided by the smallest number of mole which is 1.17

$$\text{Si: } \frac{1.17}{1.17} = 1 \qquad \text{F: } \frac{3.53}{1.17} \approx 3$$

Thus the empirical formula is SiF_3

Then we calculate the molar mass of the empirical formula

SiF_3= 85.09 g/mol

$$\text{Ratio} = \frac{\text{molar mass of compound}}{\text{empirical molar mass}}$$

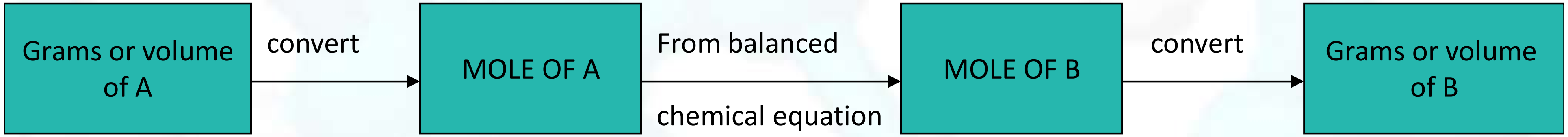
$$\text{Ratio} = \frac{168.5}{85.09} \approx 2$$

Molecular formula = empirical formula x ratio

= SiF_3 x 2 = Si_2F_6

Gas Stoichiometry

- In chapter 3 we learned how to calculate the product amount if we know the amount of reactant or how to calculate the amount of reactant if know the amount of product.
- The relationship was between n and m.
- In gases we can do the same however the relationship is between V and n.



Grams or volume of A → convert → MOLE OF A → From balanced chemical equation → MOLE OF B → convert → Grams or volume of B

We can use the volume only when the product and reactant are gases. And when T and P are constant

Gas Stoichiometry

Example 1:

Calculate the volume of O_2 (in L) required for the complete combustion of 7.64 L of C_2H_2 measured at the same temperature and pressure.?

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

From equation

2 mol $C_2H_2 \longrightarrow$ 5 mol O_2

2L $C_2H_2 \longrightarrow$ 5L O_2

7.64 L $C_2H_2 \xrightarrow{X}$? L O_2

5 x 7.64 = 2 x ?

Volume of O_2 = 5 X 7.64 / 2 = 19.1 L

Gas Stoichiometry

Example 2:

Sodium azide (NaN_3) is used in some automobile air bags. The impact of a collision triggers the decomposition of NaN_3 as follows:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

The nitrogen gas produced quickly inflates the bag between the driver and the windshield and dashboard. Calculate the volume of N_2 generated at 80 °C and 823 mmHg by the decomposition of 60g NaN_3?

T = 80 °C = 80 +273 = 353K

P= 823 mmHg

1 atm =760 mmHg

?atm = 823mmHg

823 x 1 = 760 x ? = 823/760 =1.083atm

m= 60g.

First convert g to mole

n=m/MM

= 60 / 65.02 =0.923 mol

From equation

2 mol NaN_3 $\longrightarrow$ 3 mol N_2

X

0.923 mol NaN_3 $\longrightarrow$? mol N_2

3 x 0.923 = 2 x ?

Mole of N_2= 3 x 0.923 / 2

Mole of N_2 = 1.38mol

Gas Stoichiometry

PV = nRT

$$V = \frac{n\,R\,T}{P}$$

$$V = \frac{1.38 \text{ x } 0.0821 \text{ x } 353}{1.083}$$

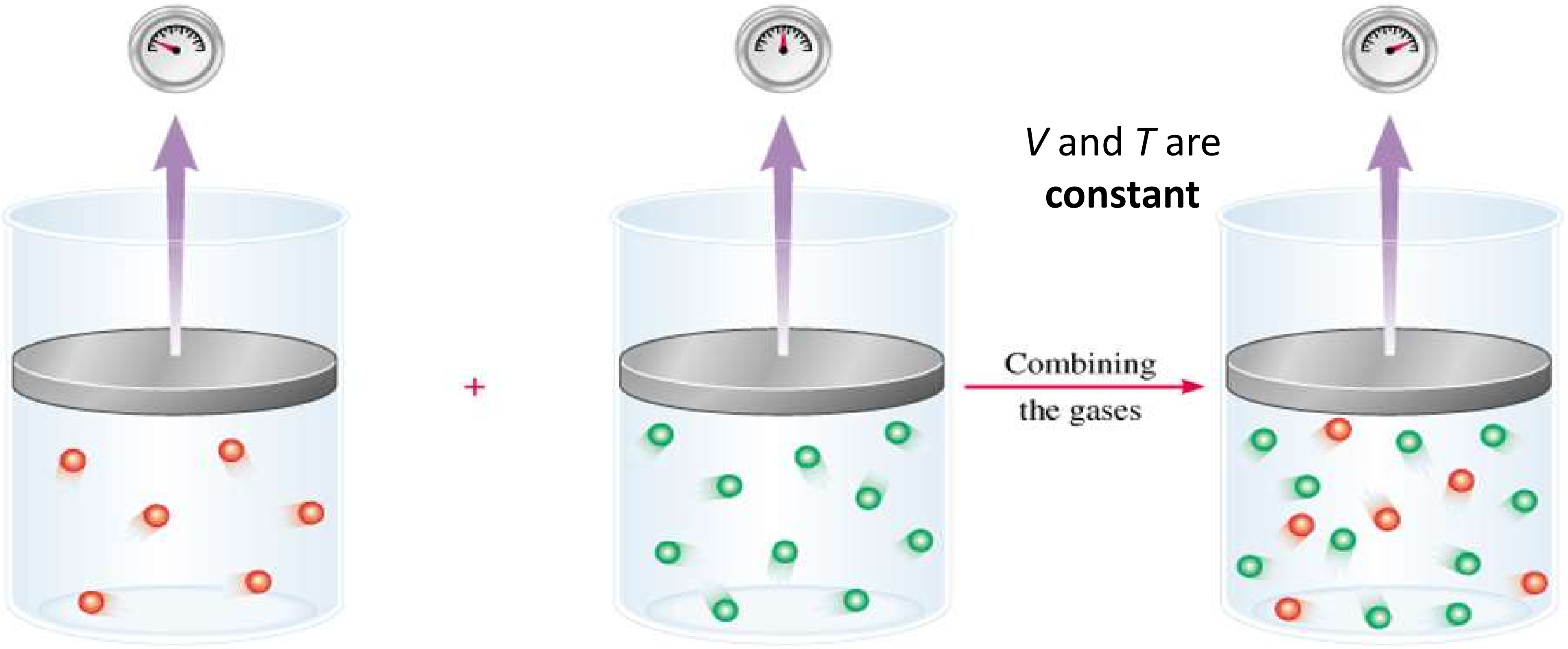
V = 36.9L

Dalton’s Law of Partial Pressure

- If we have a mixture of gases the total pressure is related to partial pressures of each gas.
- Partial pressure is the pressures of individual gas components in the mixture of gases.

Dalton’s law of partial pressures: the total pressure of a mixture of gases is just the sum of the pressures that each gas would exert if it were present alone.

Dalton’s Law of Partial Pressure



V and *T* are **constant**

P_1

P_2

$P_{total} = P_1 + P_2$

Dalton's Law of Partial Pressure

$$P_A = \frac{n_A RT}{V}$$

$$P_B = \frac{n_B RT}{V}$$

$P_T = P_A + P_B$

$$P_T = \frac{n_A RT}{V} + \frac{n_B RT}{V}$$

$$P_T = \frac{RT}{V}(n_A + n_B)$$

$$P_T = \frac{n R T}{V}$$

Where $n = n_A + n_B$

Dalton's Law of Partial Pressure

- Mole fraction: is a dimensionless quantity that expresses the ratio of the number of moles of one component to the number of moles of all components present.

$P_A = n_A RT/V$ Divided by P_T

$$\frac{P_A}{P_T} = \frac{n_A \cancel{RT/V}}{(n_A + n_B)\cancel{RT/V}}$$

$$\frac{P_A}{P_T} = \frac{n_A}{n_A + n_B}$$

$$= X_A$$

$$\boxed{X_i = \frac{n_i}{n_T}} \qquad \frac{P_i}{P_T} = X_i \qquad \boxed{P_i = X_i P_T}$$

If we have gas mixture consist of two gases (A and B)

Then the sum of all mole fraction for the same mixture is 1

$$\boxed{X_A + X_B = \frac{n_A}{n_B + n_A} + \frac{n_B}{n_A + n_B} = 1}$$

Dalton’s Law of Partial Pressure

Example:

A mixture of gasses contains 4.46 moles of Ne, 0.74 mole of Ar, and 2.15 moles of Xe. Calculate the partial pressures of the gases if the total pressure is 2.00 atm at a certain temperature.?

First we have to determine the molar fraction of each gas

$$X_i = \frac{n_i}{n_T}$$

$$X_{Ne} = \frac{4.46}{4.46 + 0.74 + 2.15} = 0.607$$

$$X_{Ar} = \frac{0.74}{4.46 + 0.74 + 2.15} = 0.1$$

$$X_{Xe} = \frac{2.15}{4.46 + 0.74 + 2.15} = 0.293$$

$$P_{Ne} = X_{Ne} P_T = 0.607x2 = 1.214atm$$

$$P_{Ar} = X_{Ar} P_T = 0.1x2 = 0.2atm$$

$$P_{Xe} = X_{Xe} P_T = 0.293x2 = 0.586atm$$

Dalton’s Law of Partial Pressure

- We can use Dalton’s Law of partial pressure to calculate the volume of gases over water.
- We can use this method as long as the gas we are collecting do not react with water and do not dissolve in water.
- Therefore this method is suitable for oxygen but not suitable for NH_3.
- We need to remember that above water there is always water vapor which have pressure
- Therefore if we want to collect oxygen gas:

$$P_T = P_{O_2} + P_{H_2O}$$

Dalton’s Law of Partial Pressure

Example:

Oxygen gas generated by the decomposition of potassium chlorate is collected. The volume of oxygen collected at 24 °C and atmospheric pressure of 762 mmHg is 128 ml. Calculate the mass (in g) of oxygen gas obtained. The pressure of the water vapour at 24 °C is 22.4 mmHg.?

T=24 °C = 24 + 273 = 297 K, V_{O2} =128ml = 128/1000=0.128L , m = ??

P_T = 762 mmHg,

1 atm = 760 mmHg

?atm = 762 mmHg

762 x 1= ?x 760

762 /760 =1.002atm

P_{H20}=22.4mmHg,

1 atm = 760 mmHg

?atm = 22.4 mmHg

22.4 x 1= ?x 760

22.4 /760 =0.029atm

Dalton’s Law of Partial Pressure

PV = nRT
n = PV/RT

$P_T= P_{O2} + P_{H20}$

$P_{O2}=P_T – P_{H2O}$

= 1.002 – 0.029 =0.973 atm

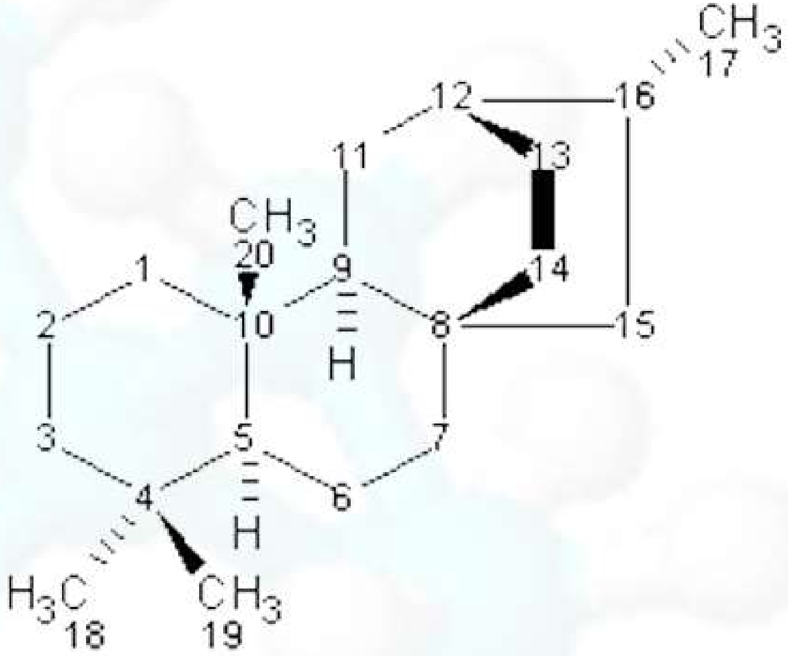
$$n = \frac{0.973x0.128}{0.0821x297} = 0.005mol$$

Mass =n x molar mass

= 0.005 x 32

=0.16 g

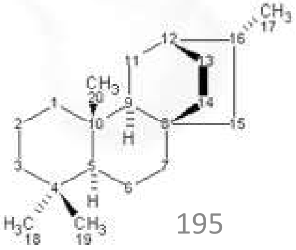
Thank you



Chapter Seven

Quantum Theory and the Electronic Structure of Atoms

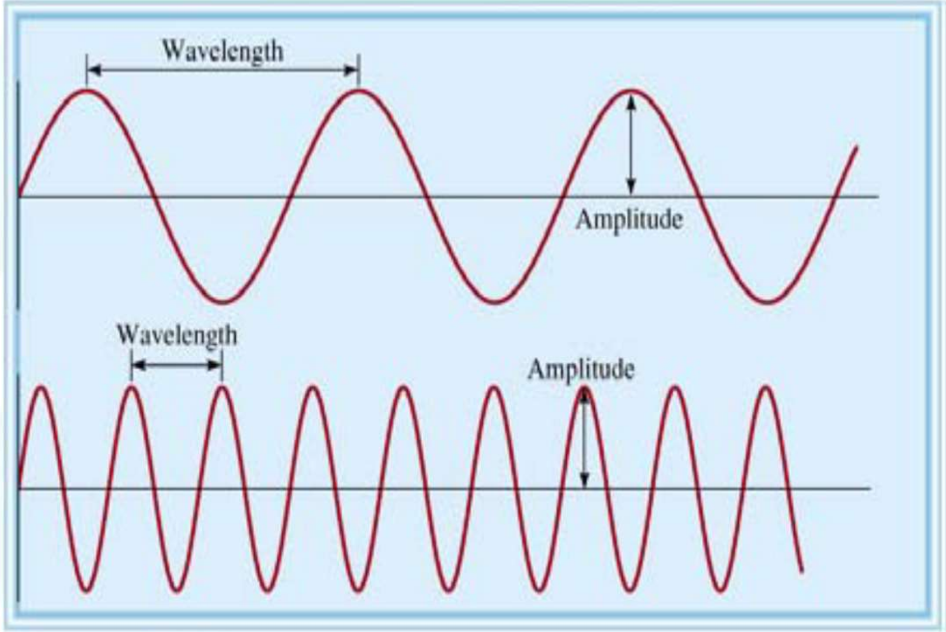
- **Chapter 7: Quantum Theory and the Electronic Structure of Atoms (268 -272 - 274-279 - 286-301)**
- 7.1 From Classical Physics to Quantum Theory
- 7.2 Bohr’s Theory of the Hydrogen Atom
- 7.4 The Dual Nature of the Electron
- 7.6 Quantum Numbers and Electromagnetic Radiation
- 7.7 Atomic Orbitals
- 7.8 Electron Configuration
- 7.9 The Building-up Principle



From Classical Physics to Quantum Theory

We have to understand something about the nature of wave before talking about Quantum theory.

- Wave: a vibrating disturbance by which energy is transmitted.
- Wavelength (λ) lambda: is the distance between identical points on successive waves.
- Frequency (ν) nu: is the number of waves that pass through a particular point in 1 second.
- Amplitude : is the vertical distance from the midline of a wave to the peak.



From Classical Physics to Quantum Theory

- Wave speed (u): depend on type of wave and the nature of the medium through which the wave is traveling.

$$\boxed{u = \lambda\nu}$$

- Wavelength usually expressed in units of meter, centimeter, or nanometer.
- Frequency is measured in hertz (Hz)

1 Hz = 1 cycle/s

Normally the word cycle is left out and we say and we expressed frequency as for example 25/s

From Classical Physics to Quantum Theory

Example:

The wavelength of the green light from a traffic signal is centered at 522 nm. What is the frequency of this radiation?

$$u = \lambda \nu$$

$$\nu = u/\lambda$$

The speed of light is known as 3 x 10^8 m/s

Because the speed of light is in m we have to change the wavelength to m

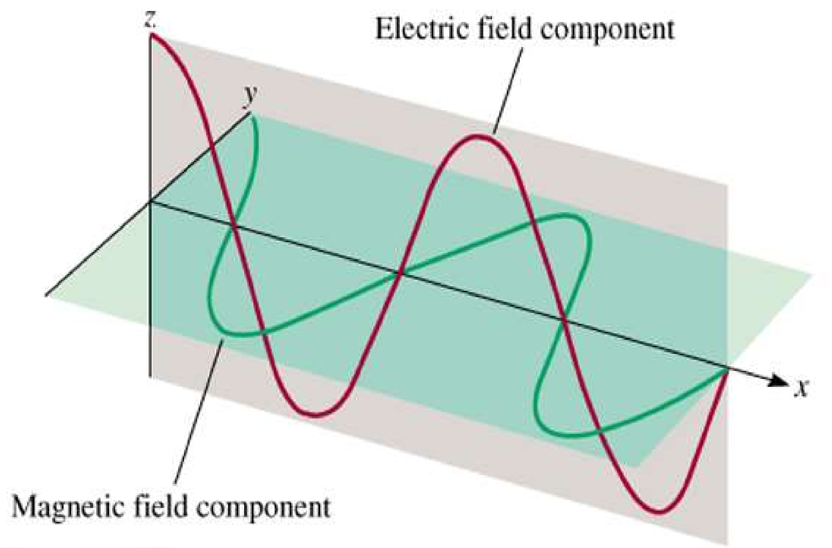
$\lambda = 522 \times 10^{-9}$ m

$$\nu = 3 \times 10^8 / 522 \times 10^{-9}$$

$$= 5.75 \times 10^{14} \text{ Hz.}$$

From Classical Physics to Quantum Theory

- There are many type of waves, such as water waves, sound waves and light waves.
- Clerk Maxwell proposed in 1873 that visible light consists of electromagnetic wave has an electric field component and a magnetic filed component. The two components have the same wavelength and frequency, and hence the speed.

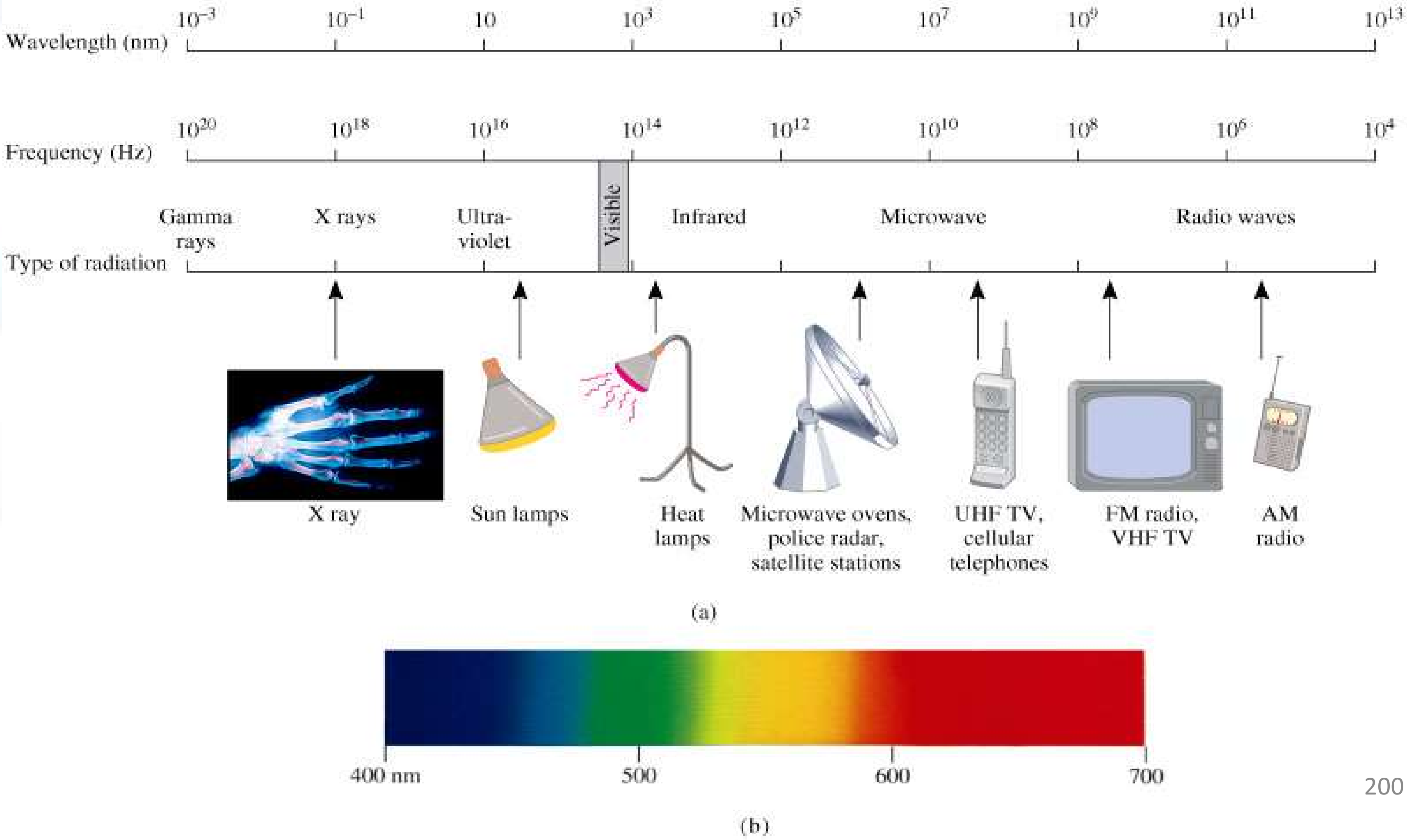


- Electromagnetic radiation is the emission and transmission of energy in the form of electromagnetic waves.
- for all electromagnetic radiation

$$c = \lambda\nu$$

Where c is the speed of light = 3 x 10^8 m/s

From Classical Physics to Quantum Theory



From Classical Physics to Quantum Theory

Example:

A photon has a frequency of 6.0 x 10^4 Hz. Convert this frequency into wavelength (nm)?

$$c = \lambda\nu$$

$$\lambda = c/\nu$$

$$= 3 \text{ x} 10^8 / 6 \text{ x} 10^4$$

$$=5 \text{ x } 10^3 \text{ m}$$

$$= 5 \text{ x } 10^{12} \text{ nm}$$

Planck’s Quantum Theory

- When solid are heated they emit electromagnetic radiation over a wide range of wavelength. Example is the dull red glow of an electric heater and the bright white light of tungsten light bulb.
- the amount of radiant energy emitted by an object at a certain temperature depends on the wavelength
- according to Plank the atoms and molecules could emit (or adsorb) energy only in discrete quantities (quantum).
- Quantum is the smallest quantity of energy that can be emitted (or absorbed) in the form of electromagnetic radiation.
- The energy of a signal quantum (E):

$$E = h\nu$$

$$\nu = \frac{c}{\lambda} \Rightarrow \boxed{E = h\frac{c}{\lambda}}$$

Where h is planck’s constant

Planck's Quantum Theory

- According to the quantum theory energy is always emitted in integral multiples of hν, for example hν, 2hν, 3hν, …… but never for exmaple 1.67hν.

Example:

Calculate the energy (in J) of:

(a) a photon with a wavelength of 5.00 x 10^4 nm (IR region)
(b) a photon with a wavelength of 5.00 x 10^{-2} nm (X-ray region)

a- λ= 5 x 10^4nm = 5 x 10^4 x 10^{-9}
= 5 x 10^{-5}m

$$E = h\frac{c}{\lambda}$$

$$E = \frac{6.63x10^{-34}x3x10^8}{5x10^{-5}} = 3.98x10^{-21}J$$

b- λ= 5 x 10^{-2} nm = 5 x 10^{-2} x 10^{-9}
= 5 x 10^{-11}m

$$E = h\frac{c}{\lambda}$$

$$E = \frac{6.63x10^{-34}x3x10^8}{5x10^{-11}} = 3.98x10^{-15}J$$

Bohr's Theory of the Hydrogen Atom

Niels Bohr (1885-1962)

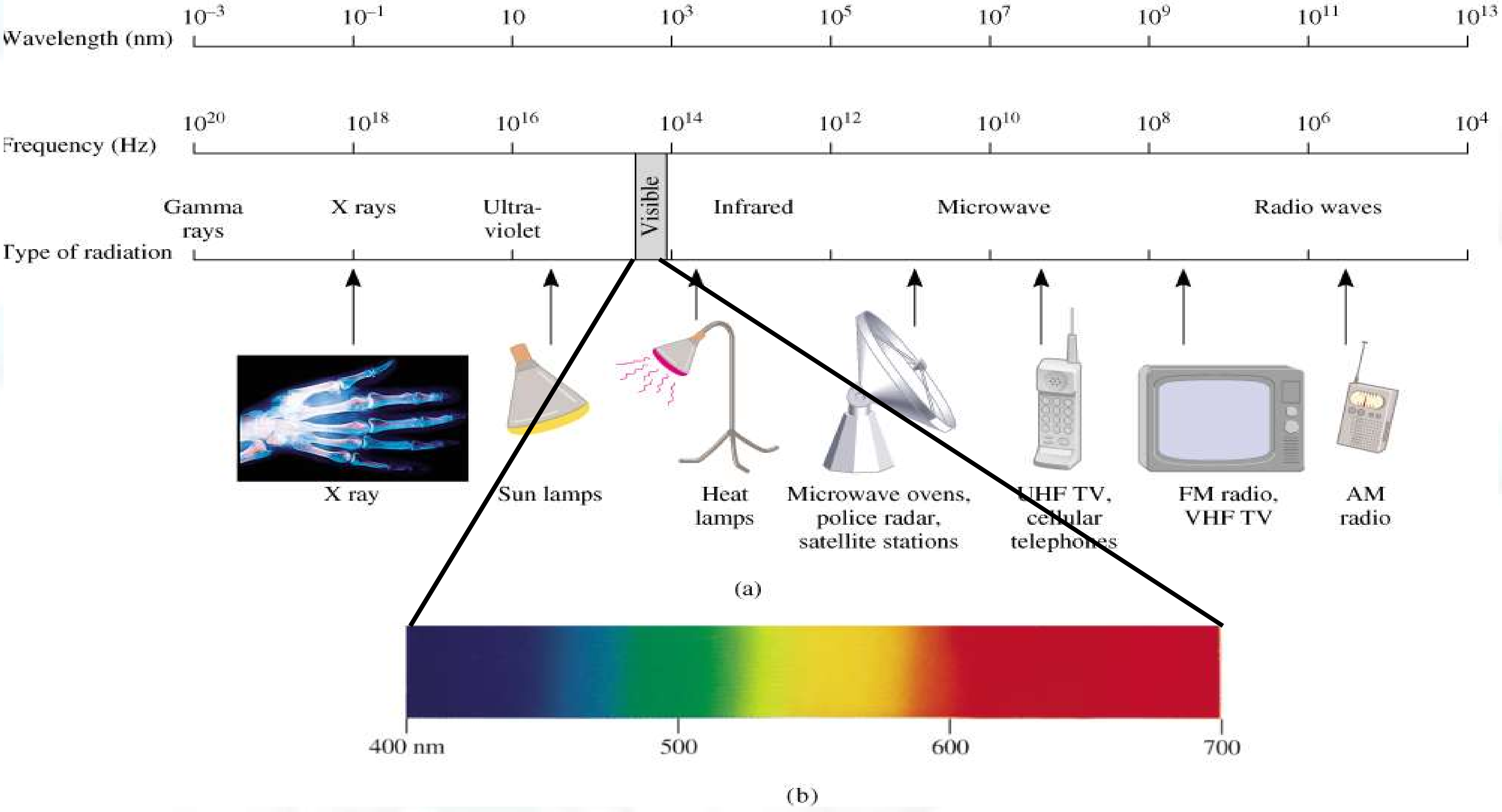
(Nobel Prize, 1922)

- Bohr's greatest contribution to science was in building a simple model of the atom.
- It was based on understanding the SHARP LINE SPECTRA of excited atoms.

Bohr’s Theory of the Hydrogen Atom

- Emission spectra is either continuous or line spectra of radiation emitted by substances.
- Continuous is a common feature to the emission spectra of the sun and of a heated solid; that is, all wavelengths of visible light are represented in the spectra.

Bohr's Theory of the Hydrogen Atom

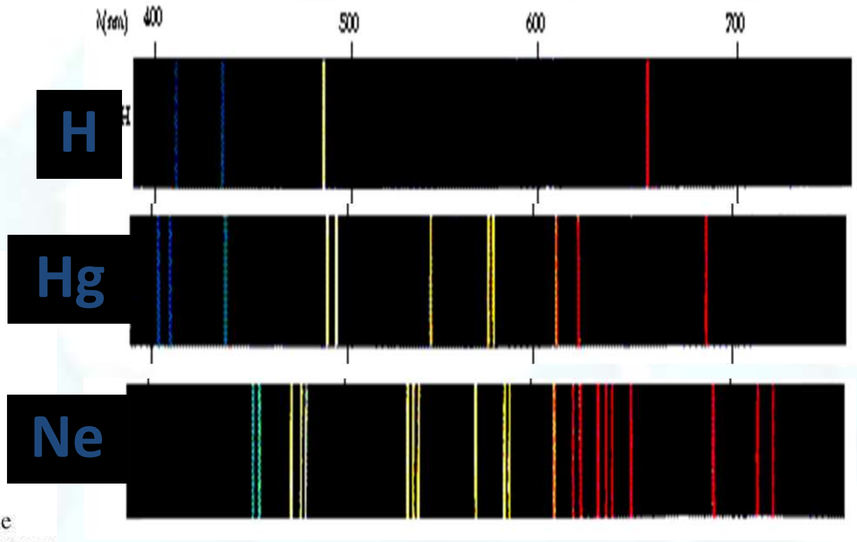


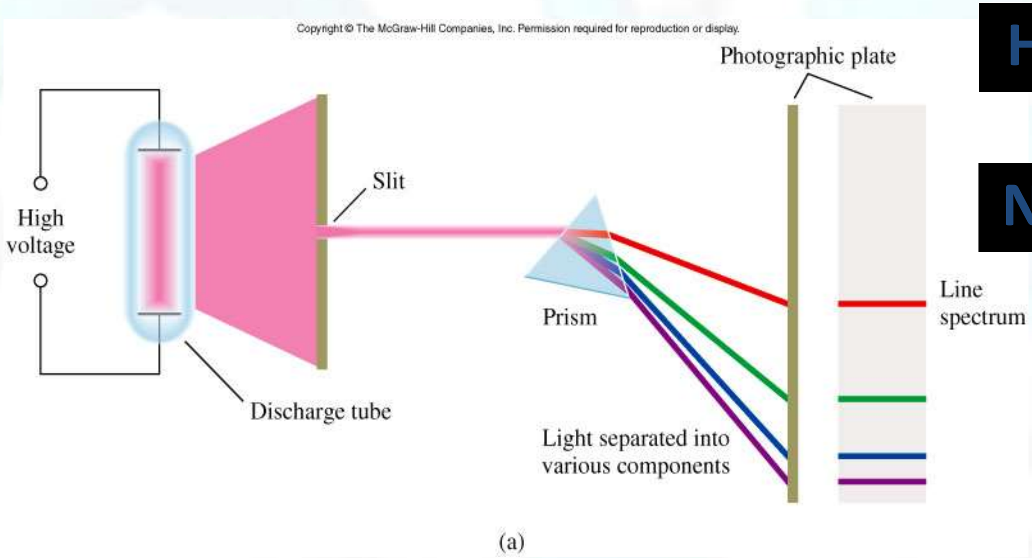
Bohr's Theory of the Hydrogen Atom

- Emission spectra is either continuous or line spectra of radiation emitted by substances.
- Continuous is a common feature to the emission spectra of the sun and of a heated solid; that is, all wavelengths of visible light are represented in the spectra.
- The emission spectra of atoms in the gas phase do not show a continuous spread of wavelengths from red to violet; rather, the atoms produce bright lines in different parts of the visible spectrum.
- Line spectra is the light emission only at specific wavelengths.
- Excited atoms emit light of only certain wavelengths
- The wavelengths of emitted light depend on the element.

Bohr’s Theory of the Hydrogen Atom

- Emission spectrum of any atom is a finger print (lines appeared at specific wavelengths)





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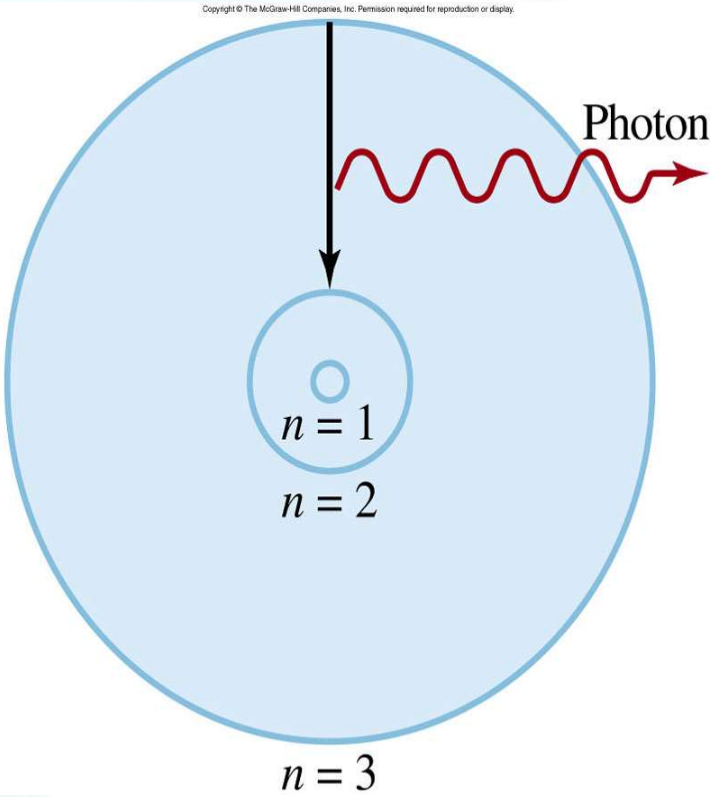
Bohr's Theory of the Hydrogen Atom

- Bohr postulated that the electron is allowed to occupy only certain orbits of specific energies and *the energies of the electron are quantized.*
- According to Bohr the emission spectrum of the H atom results from the following: the hydrogen atom is energised then electron excited to higher energy orbit and then drop to a lower-energy orbit and emitting a quantum of energy (a photon) in the form of light.



Bohr's Theory of the Hydrogen Atom

The equation that represent the energies that an electron in hydrogen atom can occupy are given by:

$$E_n = -R_H\left(\frac{1}{n^2}\right)$$

R_H is the Rodberg constant (2.18 x 10^{-18} J), n is an integer called the principle quantum number (n= 1,2,3,…..).

when n =1 this is refer to as ground state or the ground level. Which refer to the lowest energy state of a system.

n=2,3,4…. Is called an excited state or excited level, which is higher in energy than the ground state.

Bohr's Theory of the Hydrogen Atom

- The journey from a lower step to a higher step is an energy-requiring process.
- Whereas movement from a higher step to a lower step is an energy releasing process.

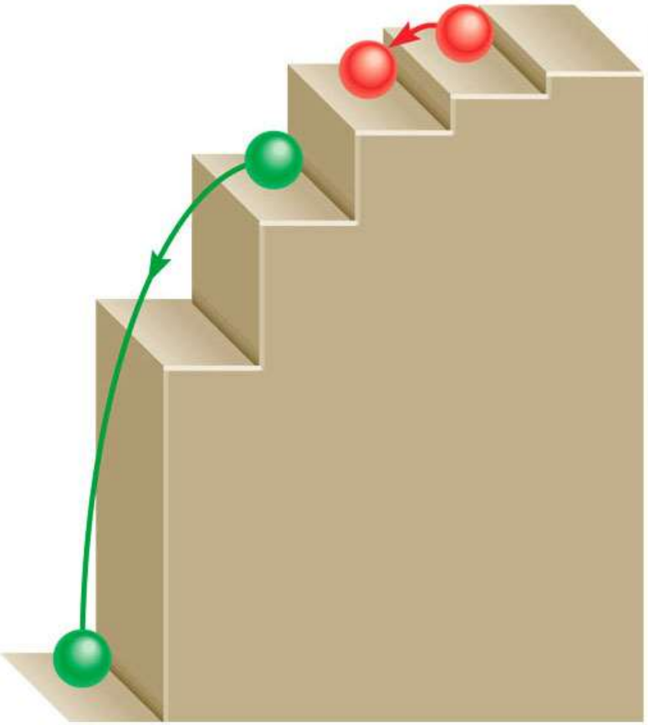
$$\Delta E = E_f - E_i$$

$$E_f = -R_H \left(\frac{1}{n_f^2}\right)$$

$$E_i = -R_H \left(\frac{1}{n_i^2}\right)$$

$$\Rightarrow \Delta E = \left(\frac{-R_H}{n_f^2}\right) - \left(\frac{-R_H}{n_i^2}\right)$$

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$



Bohr's Theory of the Hydrogen Atom

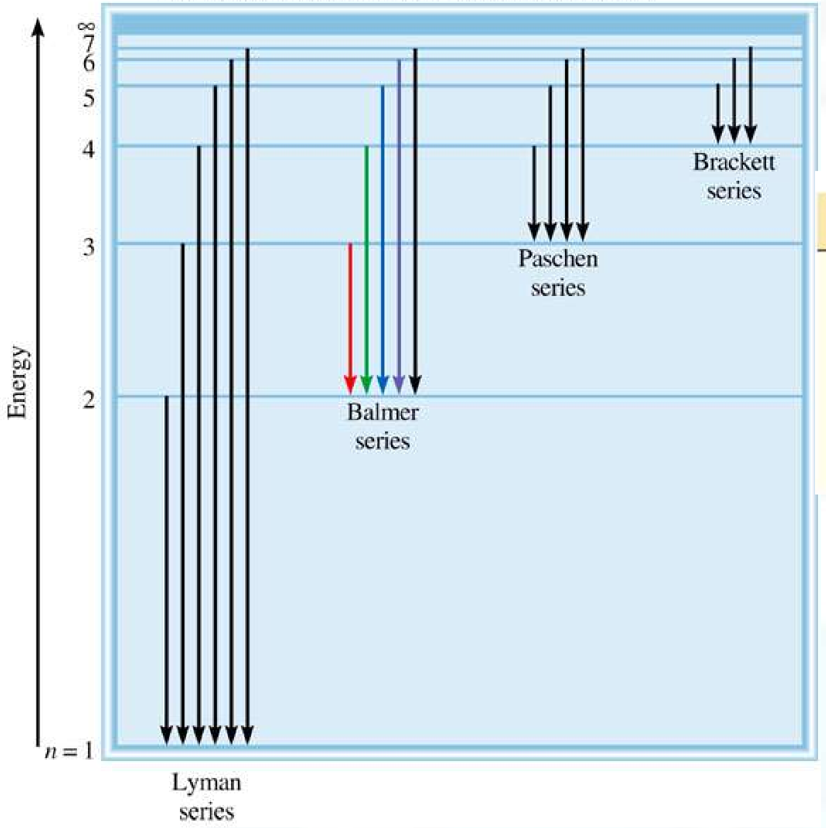
$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\Delta E = h\nu$$

$$\Delta E = h\nu = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

- If $n_f > n_i$ → (+ve) → ΔE (+ve) → Energy is absorbed
- If $n_i > n_f$ → (-ve) → ΔE (-ve) → Energy is emitted

Bohr's Theory of the Hydrogen Atom



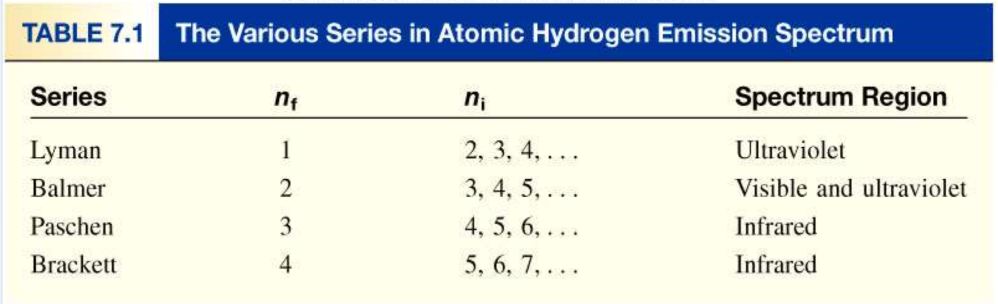


TABLE 7.1 The Various Series in Atomic Hydrogen Emission Spectrum

Series	n_f	n_i	Spectrum Region
Lyman	1	2, 3, 4, . . .	Ultraviolet
Balmer	2	3, 4, 5, . . .	Visible and ultraviolet
Paschen	3	4, 5, 6, . . .	Infrared
Brackett	4	5, 6, 7, . . .	Infrared

Bohr's Theory of the Hydrogen Atom

Example:

What is the wavelength of a photon (in nm) emitted during a transition from the $n_i = 5$ state to the $n_f = 2$ state in the hydrogen atom?

$n_i=5$, $n_f = 2$, $\lambda=?$

$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\Delta E = \frac{hc}{\lambda} \Rightarrow \lambda = \frac{hc}{\Delta E}$$

$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= 2.18x10^{-18}\ \mathrm{J}\left(\frac{1}{5^2} - \frac{1}{2^2}\right) = -4.58x10^{-19}\ \mathrm{J}$$

$$\Delta E = \frac{hc}{\lambda} \Rightarrow$$

$$\lambda = \frac{hc}{\Delta E} = \frac{6.63x10^{-34}\ \mathrm{x}\ 3.0\ \mathrm{x}10^{8}}{4.58x10^{-19}}$$

$$= 4.34x10^{-7}\ \mathrm{m} = 434\ \mathrm{nm}$$

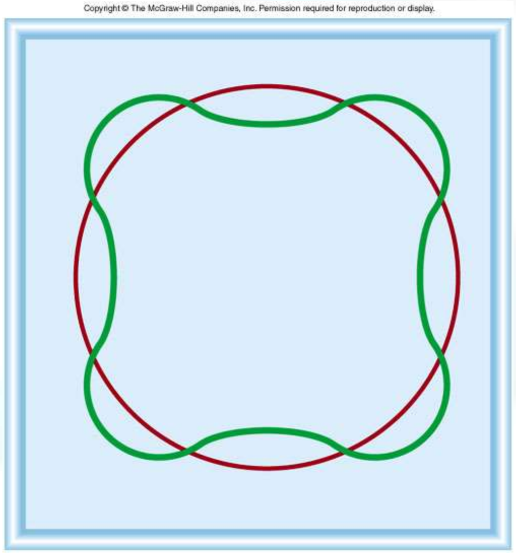
The Dual Nature of the Electron

- De Broglie suggested that particles such as electron can posses wave properties.
- According to de Borglie, an electron bond to nucleus behave like a standing wave.
- De Broglie deduced that the particle and wave

properities are related by the expression

$$\lambda = \frac{h}{mu}$$

Where λ wavelenght of moving particle, m mass (kg), u velocity of moving particle.



(a)

The Dual Nature of the Electron

Example

What is the de Broglie wavelength (in nm) associated with a 2.5 g Ping-Pong ball traveling at 15.6 m/s?

$$\lambda = \frac{h}{mu}$$

λ = ?, m= 2.5g = 2.5x10^{-3}kg, u=15.6m/s

$$\lambda = \frac{6.63x10^{-34}}{2.5x10^{-3}x15.6}$$

λ=1.7x10^{-31}m = 1.7x10^{-22}nm.

Quantum numbers

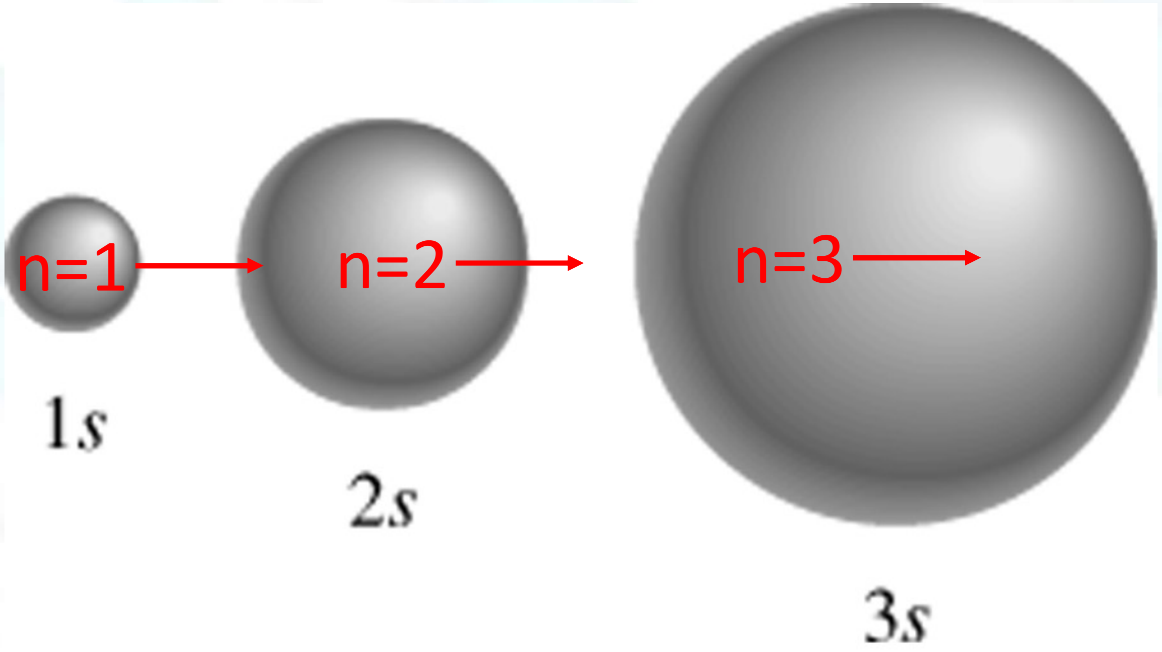
- Bohr theory did not provide a complete description of electronic behaviour in atoms and it location around the nucleus.
- In 1926 Erwin Schrödinger wrote an equation that specifies the energy states of the electron in a hydrogen atom and identifies the corresponding wave functions Ψ.
- These energy states and wave functions are characterized by a set of Quantum Numbers .
- Quantum numbers may be viewed as an electrons address.
- Schrödinger’s equation can only be solved exactly for the hydrogen atom. Must approximate its solution for multi-electron systems.
- Quantum number are a set of four values that define the energy state of an electron in an atom.

Quantum numbers

1-The Principle Quantum Number (n):

n has integer value (1,2,3,…).

n determine the energy of an orbital. Its also related to the distance between the electron and the nucleus. The larger n the grater the distance.



Quantum numbers

2- The Angular Momentum Quantum Number (l)

The value of l depend on the value of n, l= 0, (n-1).

If n=1 then l =1-1 = 0

If n =2 then l = 0 , (2-1) = 0 , 1

If n =3 then l = 0, 1 , (3-1) = 0 , 1 , 2

And so on.

l determine the shape of the orbitals.

The value of l is generally designated by the letters s, p, d, …..as follows:

l	0	1	2	3	4
Name of orbital	s	p	d	f	g

Quantum numbers

3- The Magnetic Number (m_l)

The value of m_l depend on the value of l. m_l = -l ,…….0,……+l

If l=0 then m_l =0

If l =1 then m_l = -1,0,1

If l =2 then m_l = -2,-1,0,1,2

And so on.

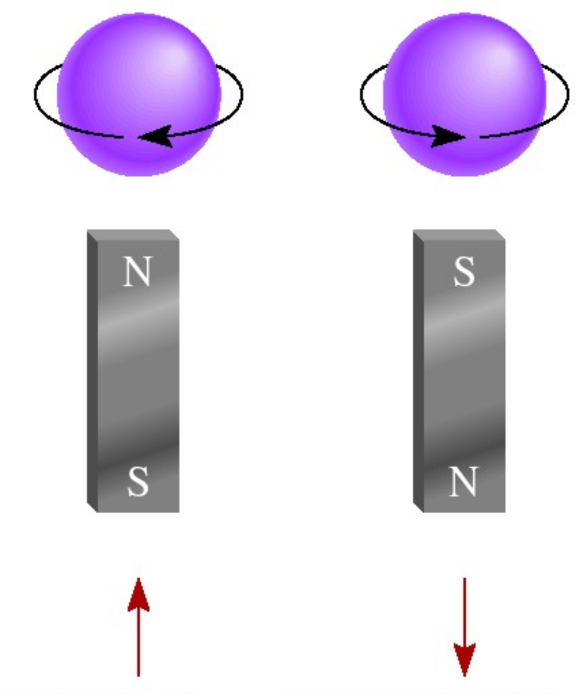
m_l describe the orientation of the orbital in space.

Quantum numbers

4- The Electron Spin Quantum Number (m_s)

m_s has two value – ½ or + ½ .

m_s determine the spin of electron.



Quantum numbers

Example :

List the values of n, l and ml for orbitals in 4d subshell? What is the total number of orbital in 4d?

n= 4

for d l=2

m_l = -2,-1,0,1,2

Number of orbital = 5

Atomic Orbitals

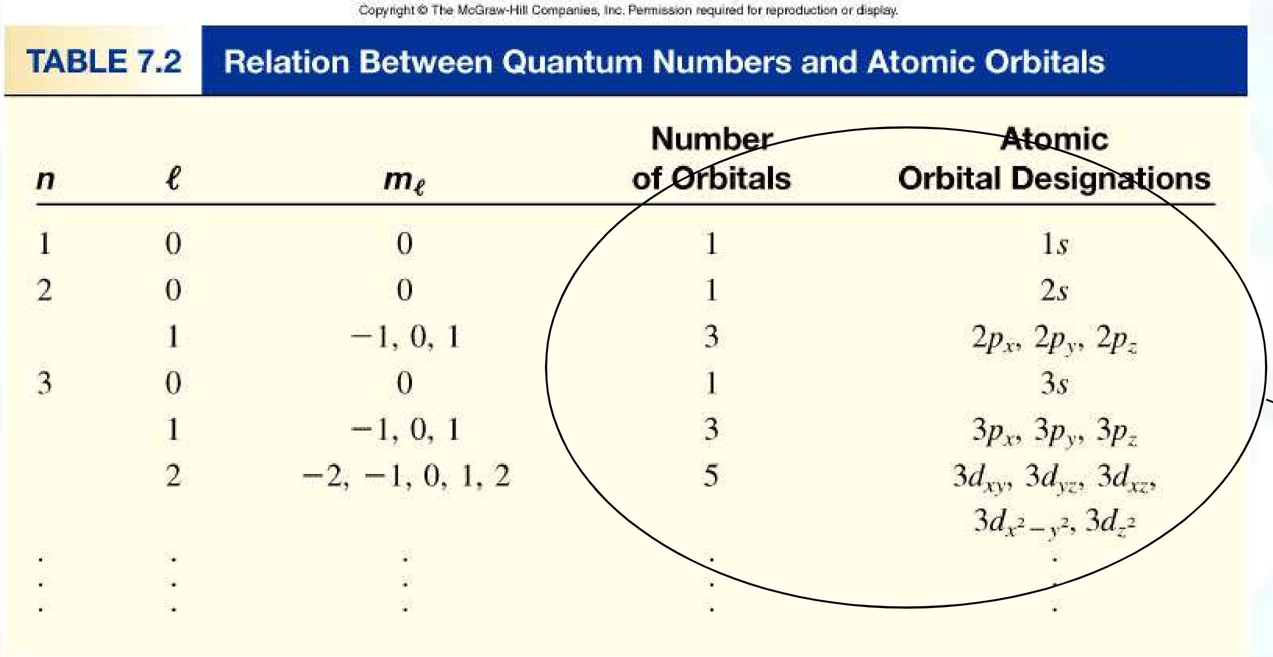


TABLE 7.2 Relation Between Quantum Numbers and Atomic Orbitals

n	ℓ	m_ℓ	Number of Orbitals	Atomic Orbital Designations
1	0	0	1	$1s$
2	0	0	1	$2s$
	1	−1, 0, 1	3	$2p_x$, $2p_y$, $2p_z$
3	0	0	1	$3s$
	1	−1, 0, 1	3	$3p_x$, $3p_y$, $3p_z$
	2	−2, −1, 0, 1, 2	5	$3d_{xy}$, $3d_{yz}$, $3d_{xz}$, $3d_{x^2-y^2}$, $3d_{z^2}$
⋮	⋮	⋮	⋮	⋮

2l +1

Atomic Orbitals

Example

What is the total number of orbitals associated with the principal quantum number n = 3 ?

$n=3$

$l=0,1,2$

1- $m_l = 0$ ======> 1 orbital

2- $m_l = -1,0,1$ =====> 3 orbital

2- $m_l = -2,-1,0,1,2$ ======> 5 orbital

Total number of orbital 9

Or (2x0 +1) + (2x1 +1) + (2x2 +1) =1 + 3 + 5 = 9

OR

Number of orbital = $n^2 = 3^2 = 9$

Atomic Orbitals

- The four quantum number for specific electron can written as (n,l,m_l,m_s).

Example :

Write the four quantum numbers for an electron in a 3p orbital?

$n=3$, $l = 1$, $m_l= -1,0,1$, $m_s= -$ ½ or ½

(3,1,-1,-1/2)

(3,1,0,-1/2)

(3,1,1,-1/2)

(3,1,-1,1/2)

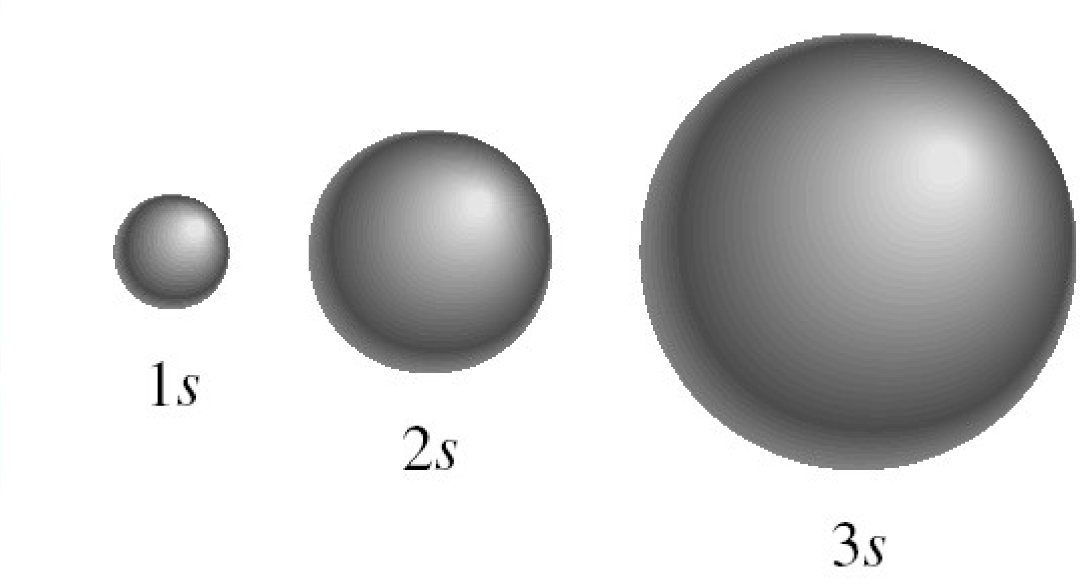
(3,1,0,1/2)

(3,1,1,1/2)

Atomic Orbitals

S orbital
l=0

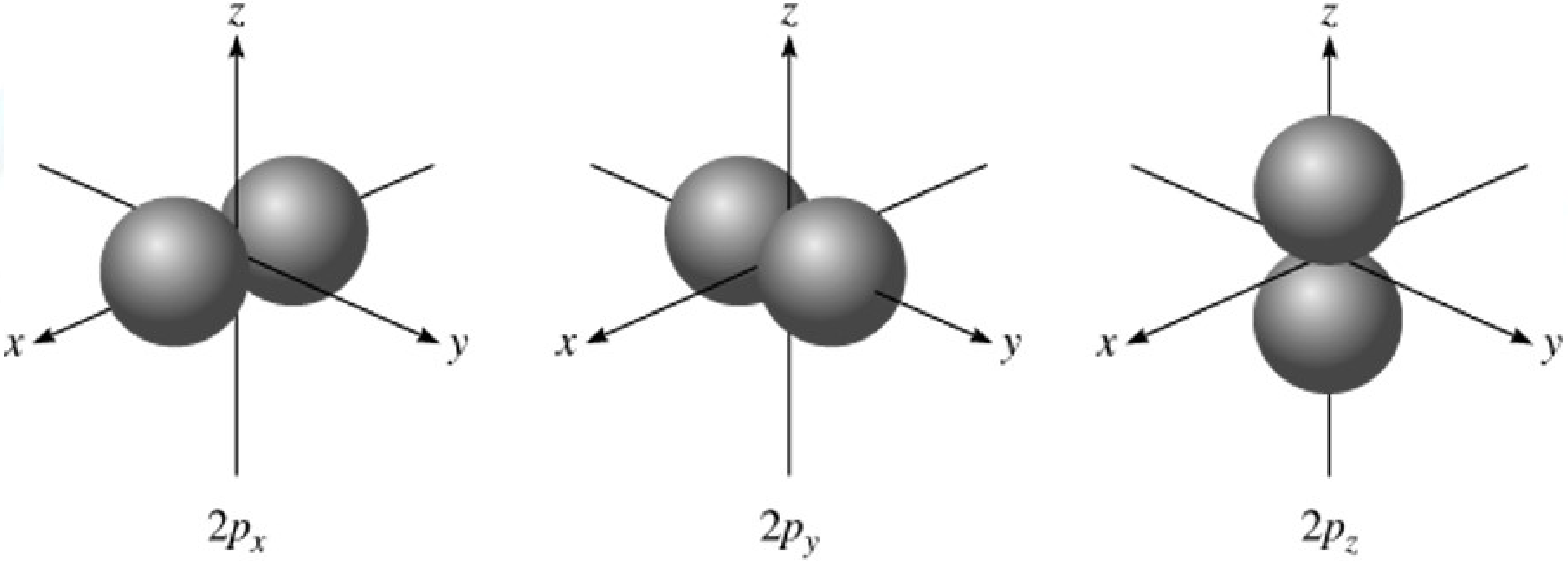
Spherical



Atomic Orbitals

P orbital
l=1

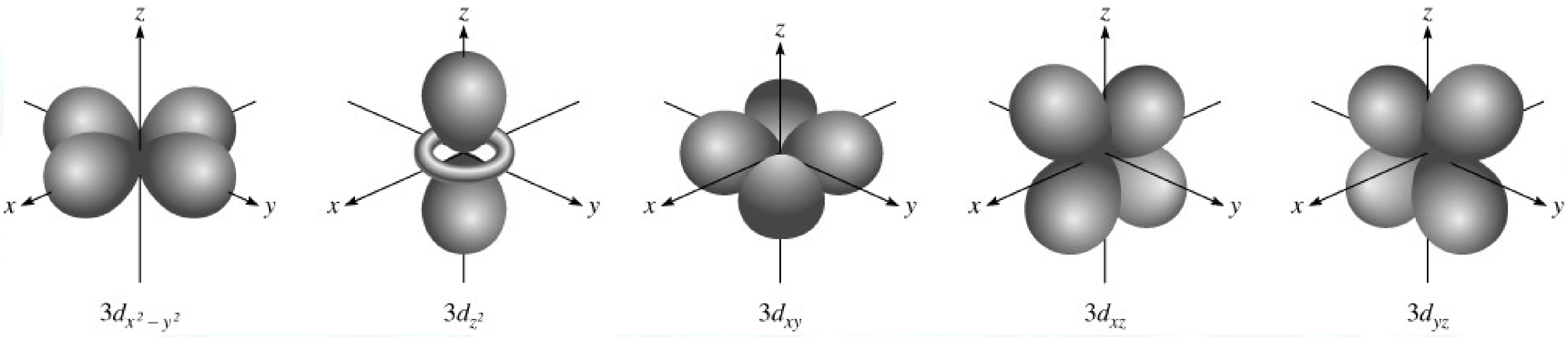
dumbbell



$m_l = -1$ $m_l = 0$ $m_l = 1$

Atomic Orbitals

d orbital
l=2

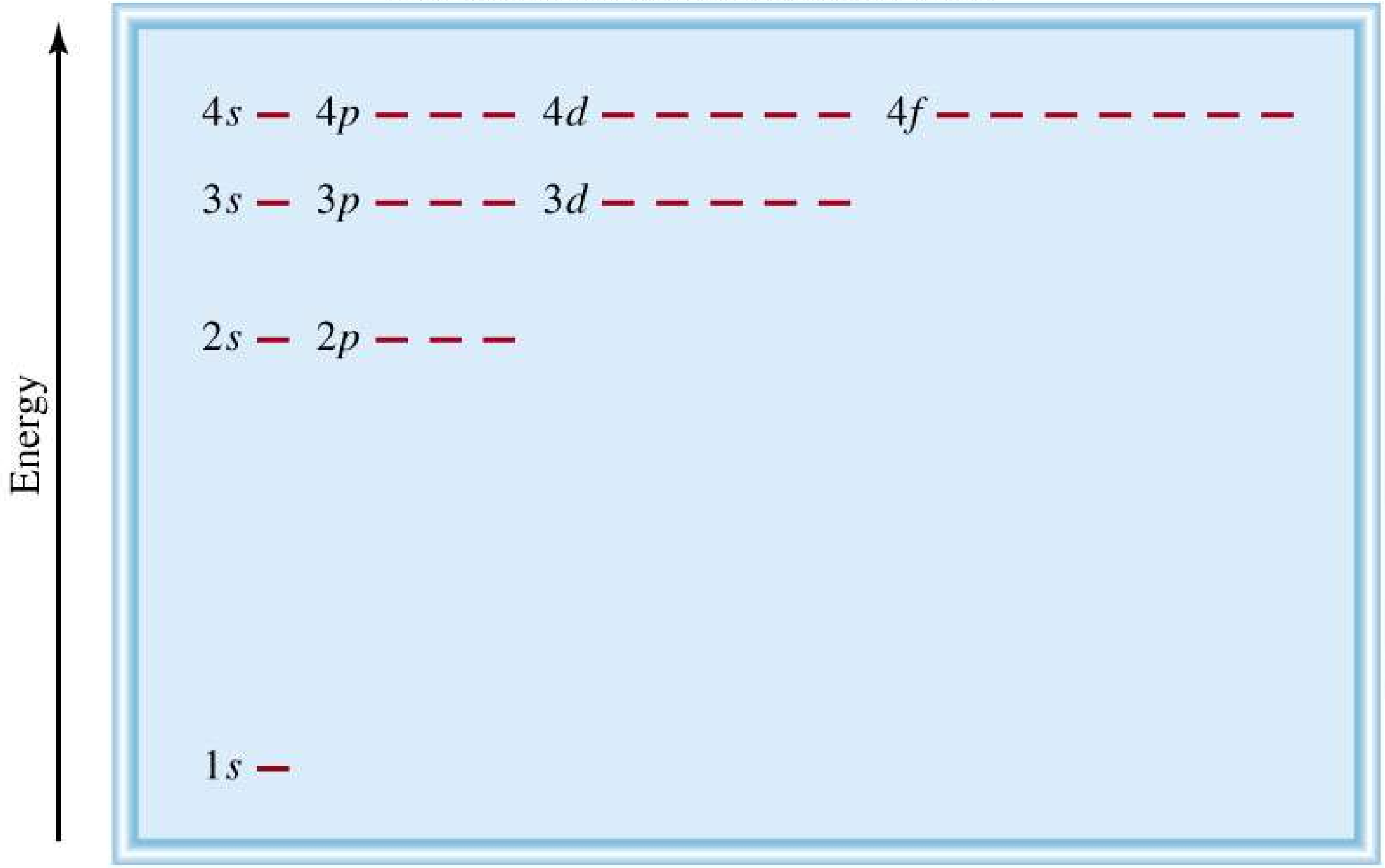


$m_l = -2$ $m_l = -1$ $m_l = 0$ $m_l = 1$ $m_l = 2$

The Energies of Orbitals

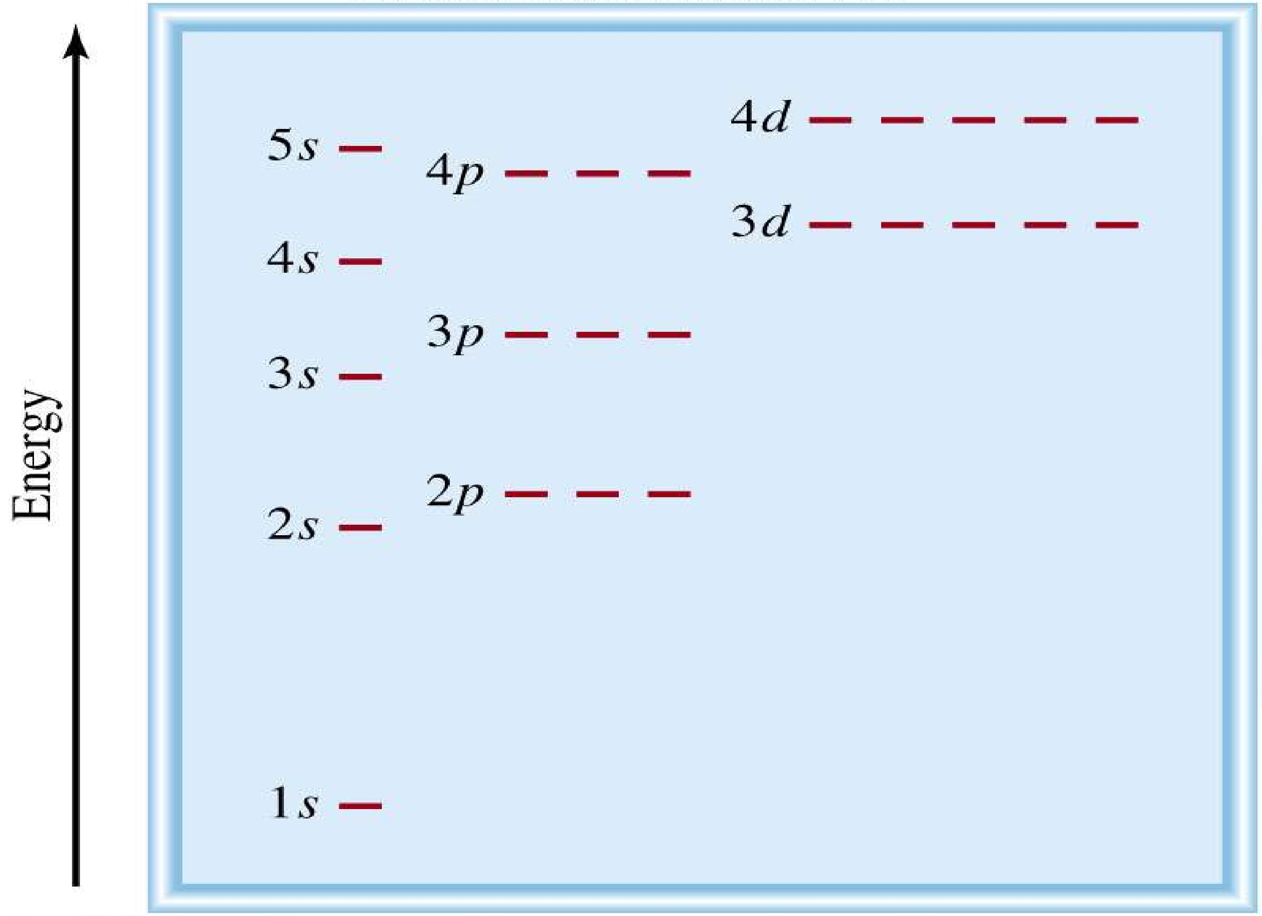
- For hydrogen atom, the energies of hydrogen orbitals increase as following:
- 1s < 2s=2p < 3s=3p=3d < 4s=4p=4d=4f <……..
- Energy only depends on principal quantum number n
- Orbitals on the same energy level have the same energy.
- For atom with multi electron, energy of orbitals depend on n and l. it follow:
- 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s

The Energies of Orbitals



For hydrogen atom

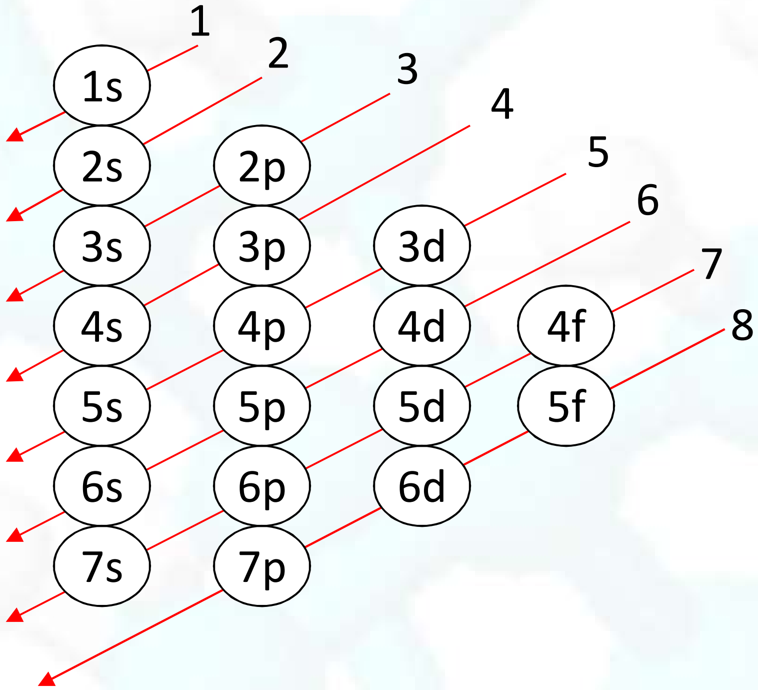
The Energies of Orbitals



For multi electron atom

The Energies of Orbitals

Order of orbitals (filling) in multi-electron atom

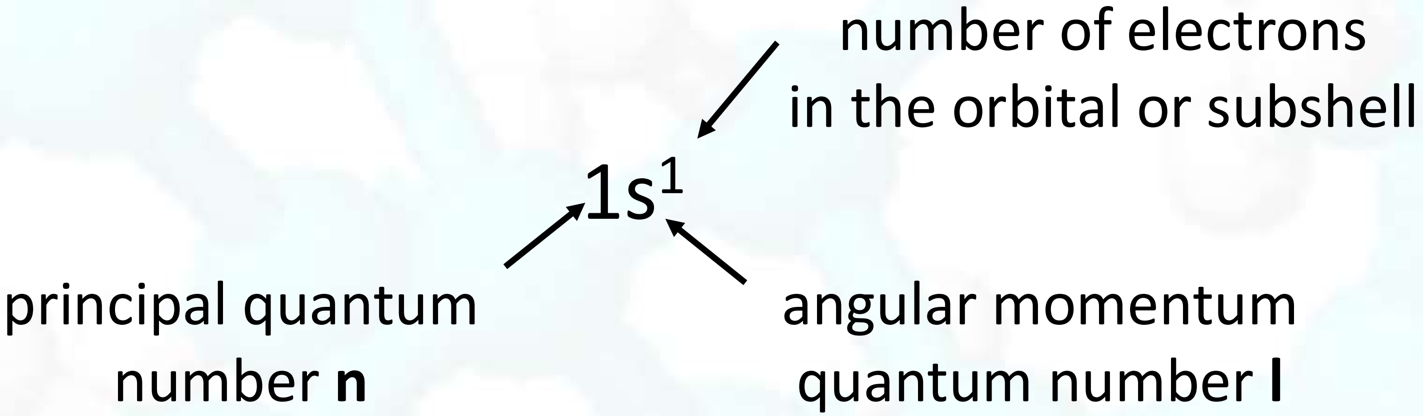


s sublevel	2 electrons
p sublevel	6 electrons
d sublevel	10 electrons
f sublevel	14 electrons

$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p$

Electron configuration

Electron configuration: is how the electrons are distributed among the various atomic orbitals in an atom.



$1s^1$

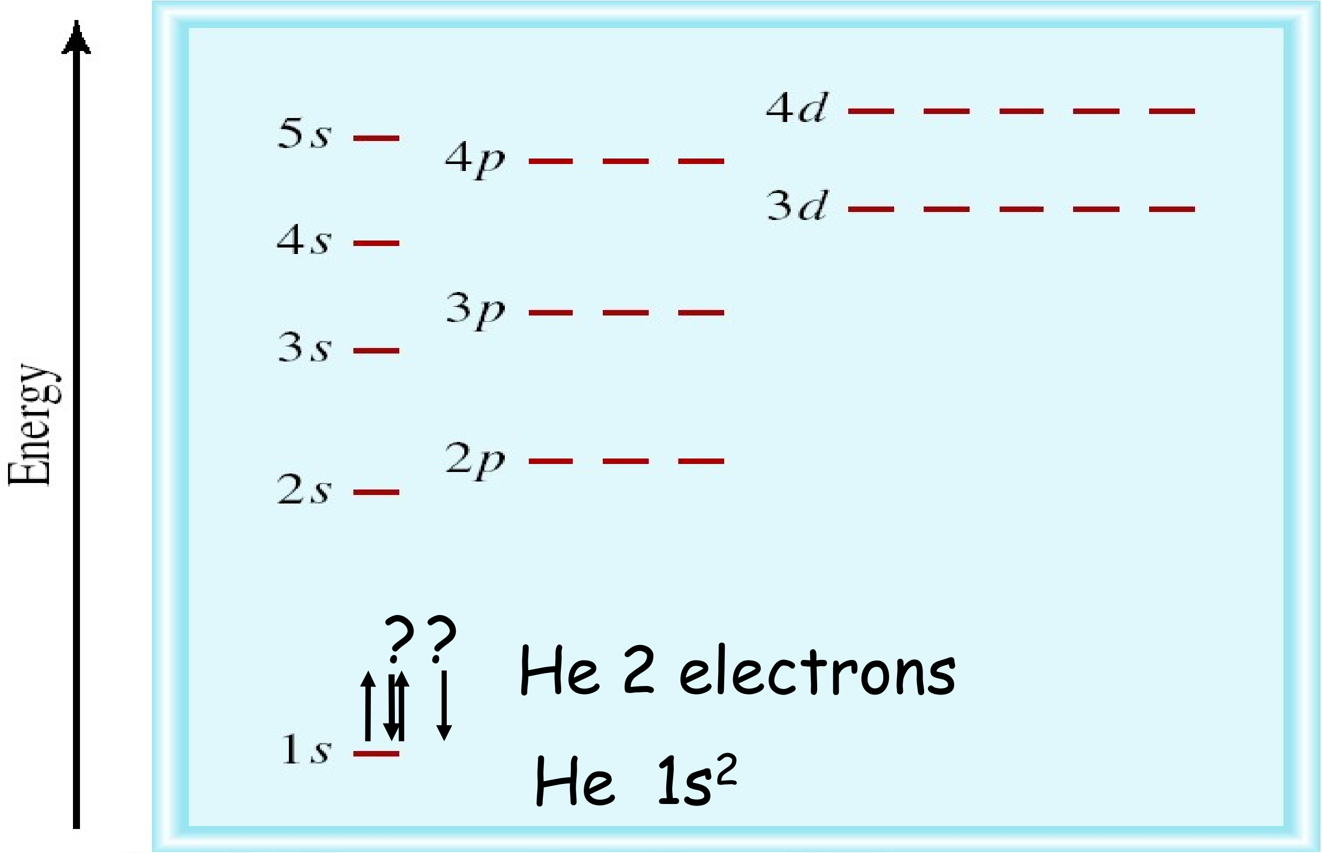
number of electrons in the orbital or subshell

principal quantum number **n**

angular momentum quantum number **l**

The Pauli Exclusion principle

NO two electrons in an atom can have the same set of four quantum numbers

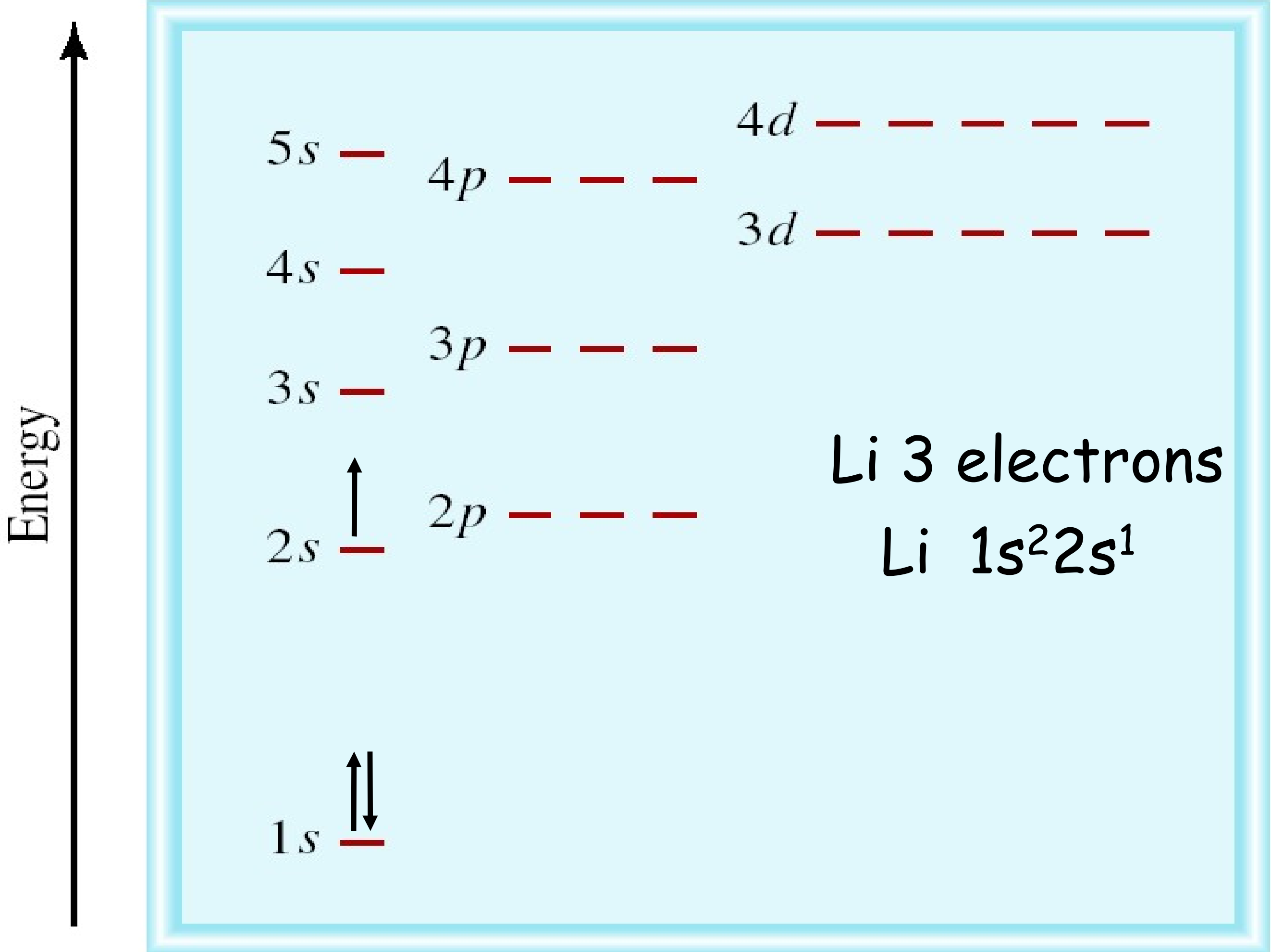


He 2 electrons

He $1s^2$

Aufbau principle

The electrons are added one by one to the atomic orbitals

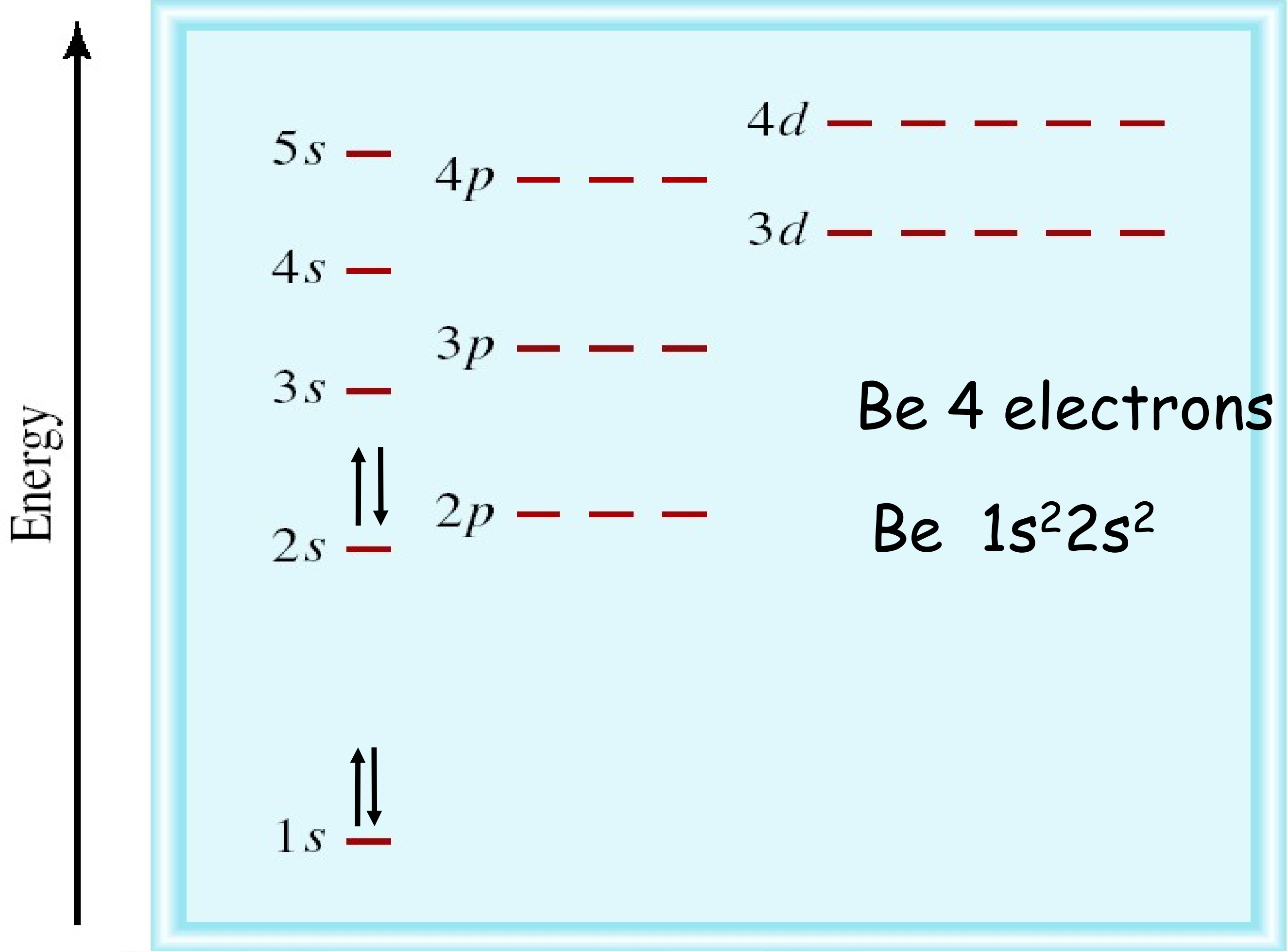


Li 3 electrons

Li $1s^2 2s^1$

Aufbau principle

The electrons are added one by one to the atomic orbitals

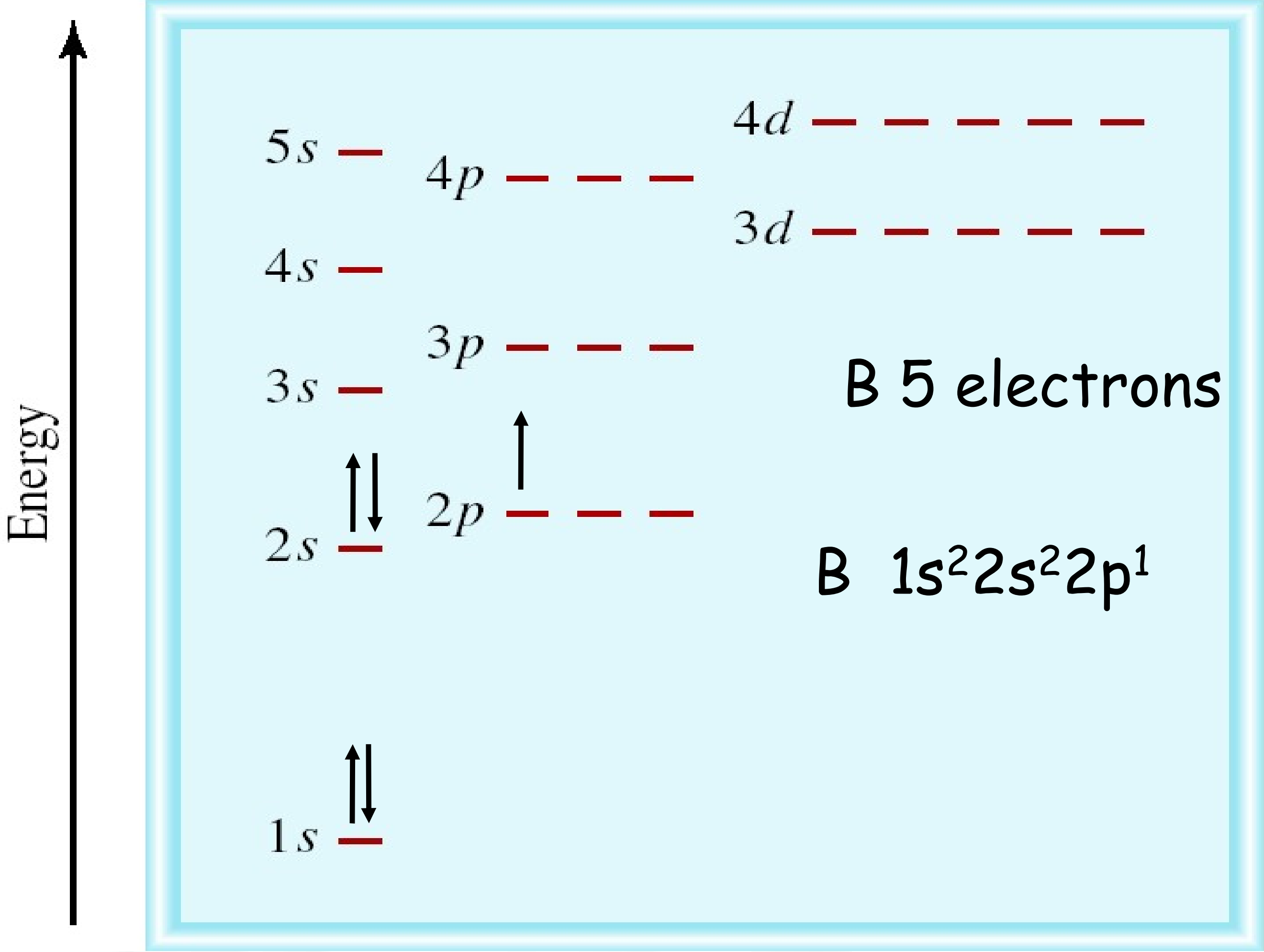


Be 4 electrons

Be $1s^2 2s^2$

Aufbau principle

The electrons are added one by one to the atomic orbitals

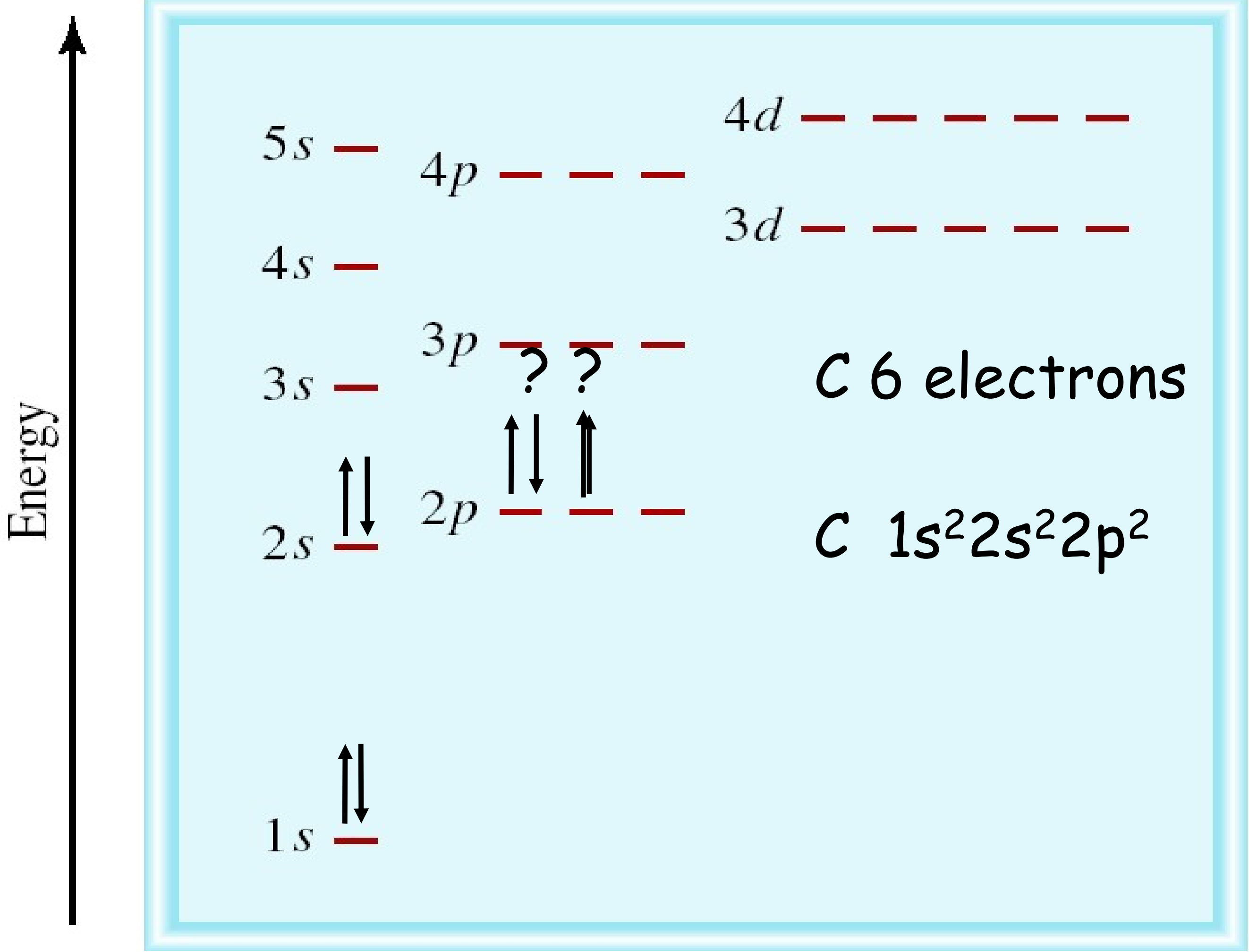


B 5 electrons

B $1s^22s^22p^1$

Hund's Rule

The most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins

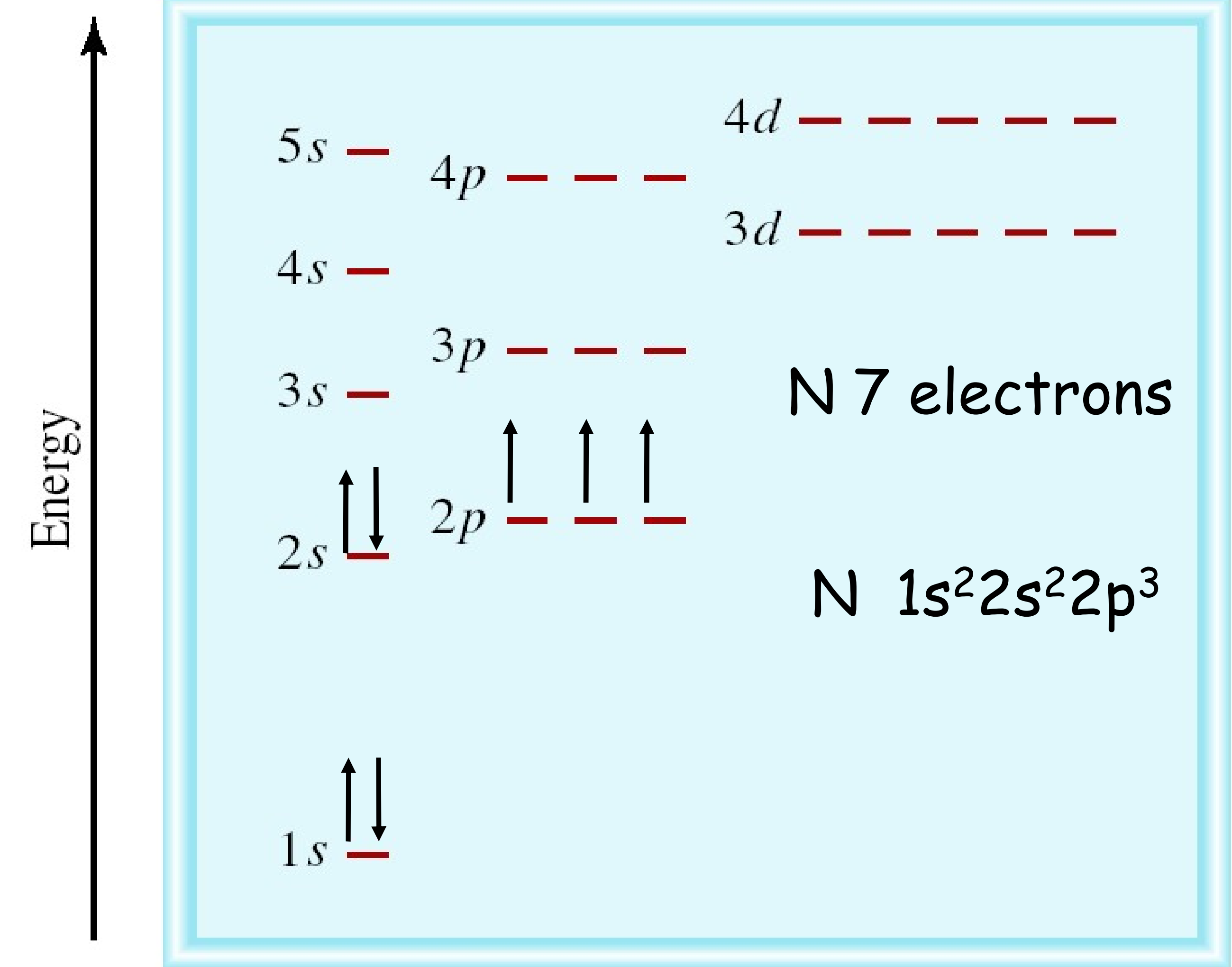


C 6 electrons

C $1s^2 2s^2 2p^2$

Hund's Rule

The most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins

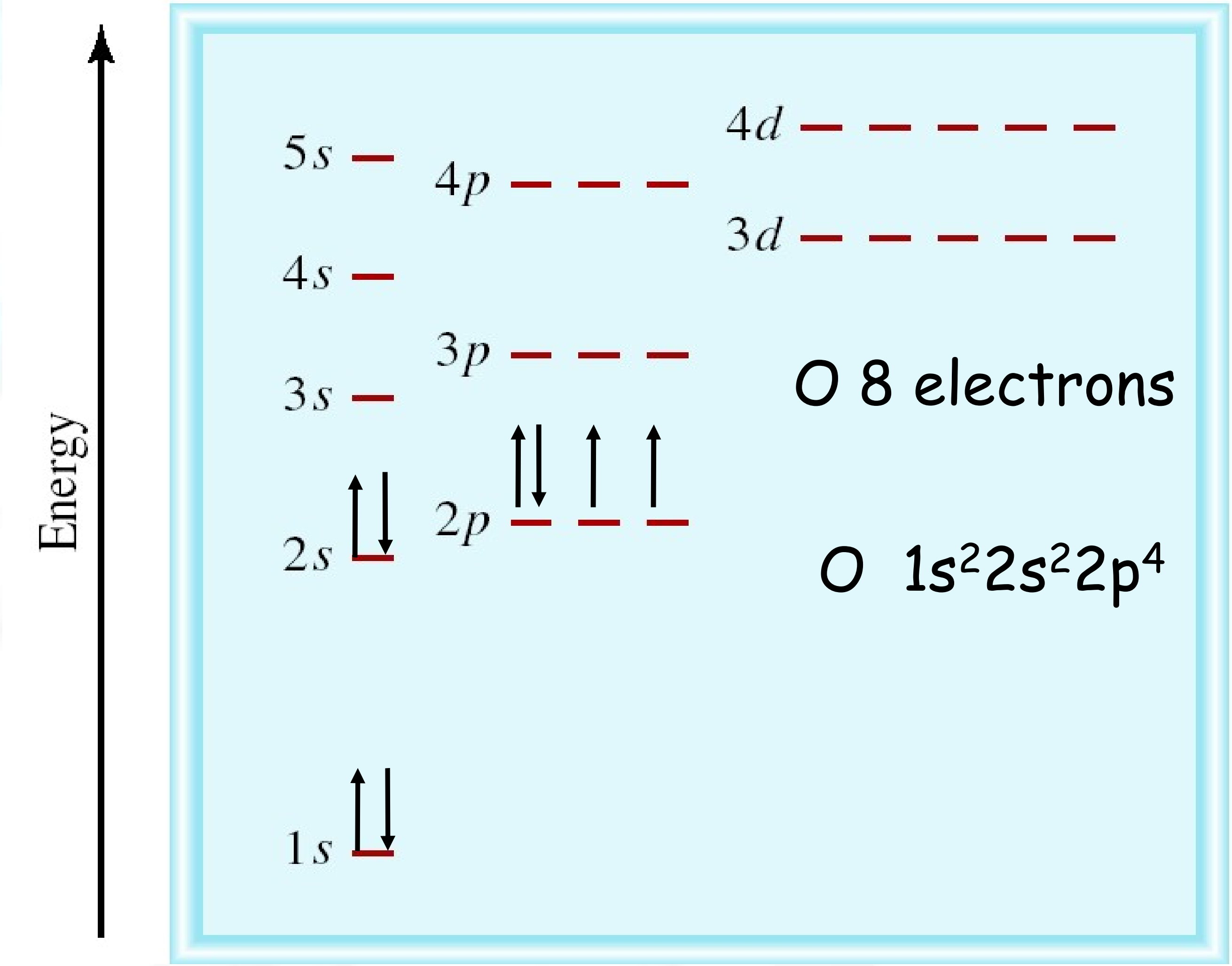


N 7 electrons

N $1s^2 2s^2 2p^3$

Aufbau principle

The electrons are added one by one to the atomic orbitals

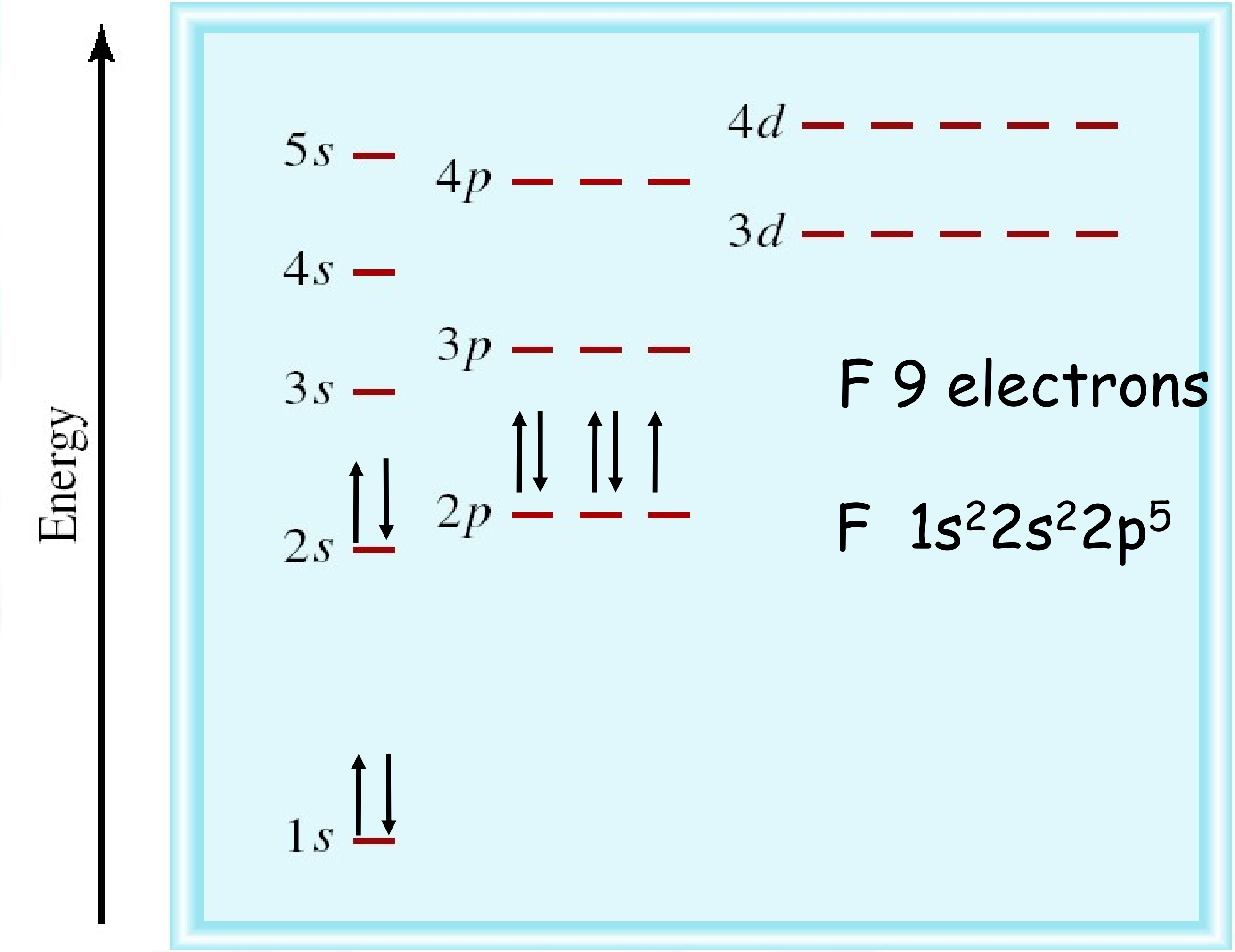


O 8 electrons

O $1s^2 2s^2 2p^4$

Aufbau principle

The electrons are added one by one to the atomic orbitals

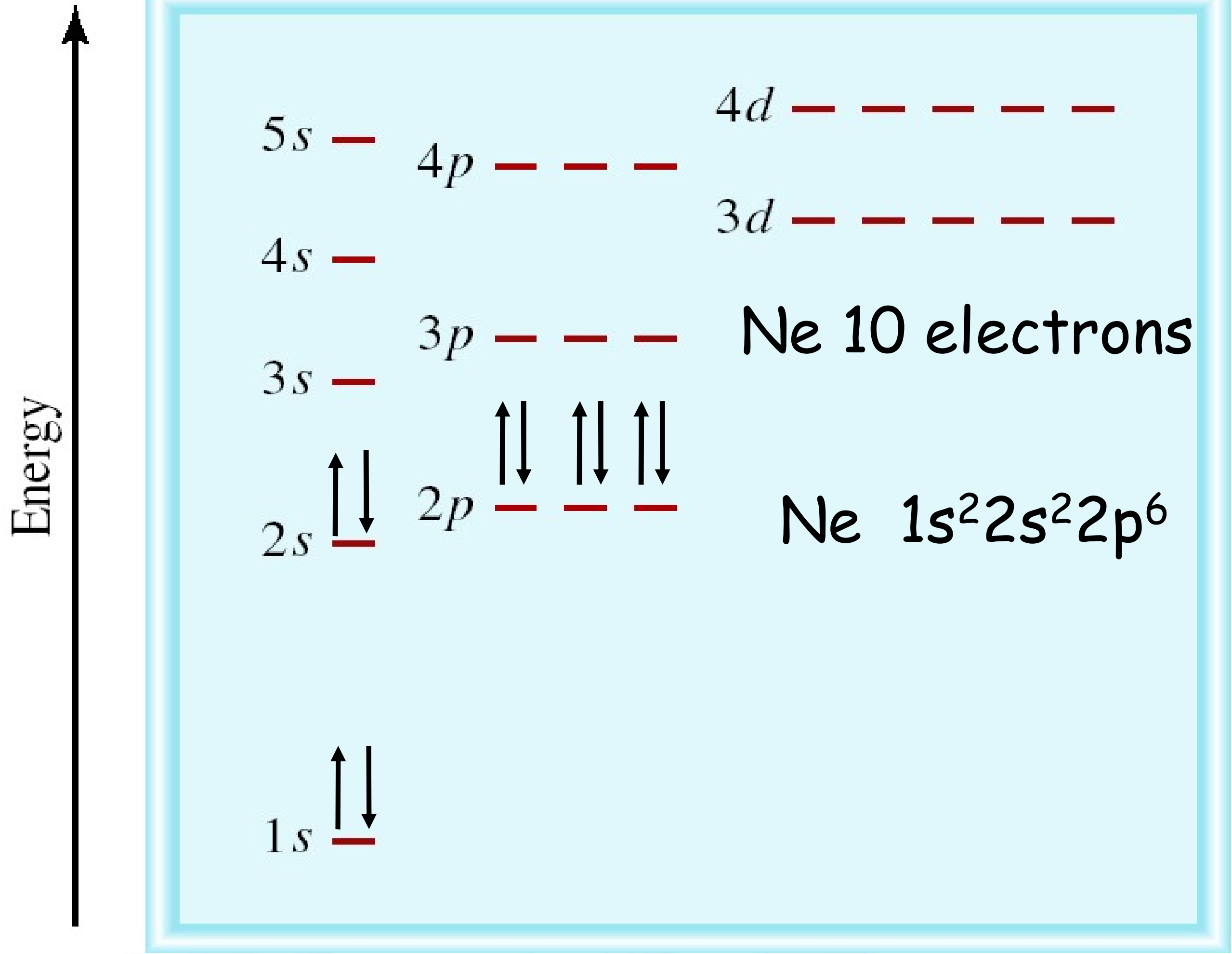


F 9 electrons

F $1s^2 2s^2 2p^5$

Aufbau principle

The electrons are added one by one to the atomic orbitals



Ne 10 electrons

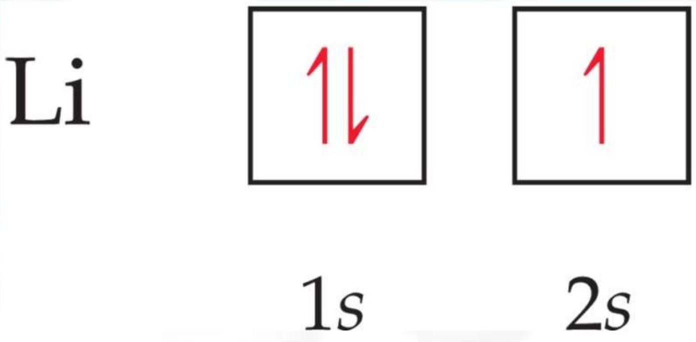
Ne $1s^2 2s^2 2p^6$

Orbital diagram

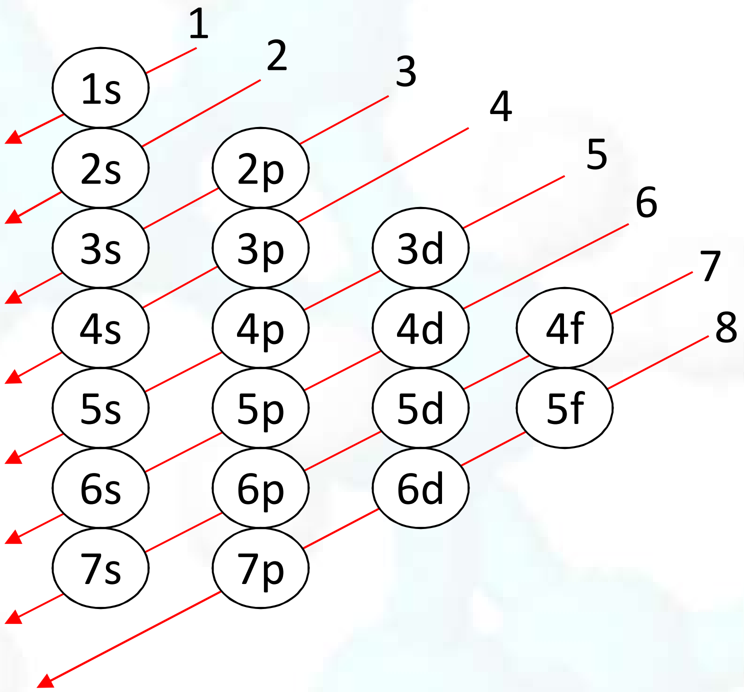
The electron configuration can also represented by the orbital diagram

In the Orbital Diagram:

- Each box represents one orbital.
- Half-arrows represent the electrons.
- The direction of the arrow represents the spin of the electron.



Orbital diagram



Example:

What is the electron configuration of Mg?

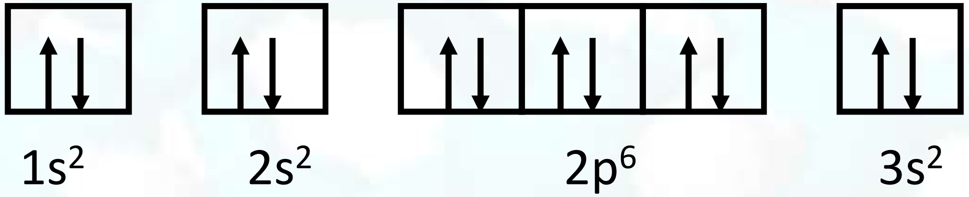
Mg atom has 12 electrons

(from Periodic table (atomic number))

The Electron Configuration: $1s^2 2s^2 2p^6 3s^2$

2 + 2 + 6 + 2 = 12 electrons

Orbital Diagram:



Orbital diagram

What is the electron configuration of K?

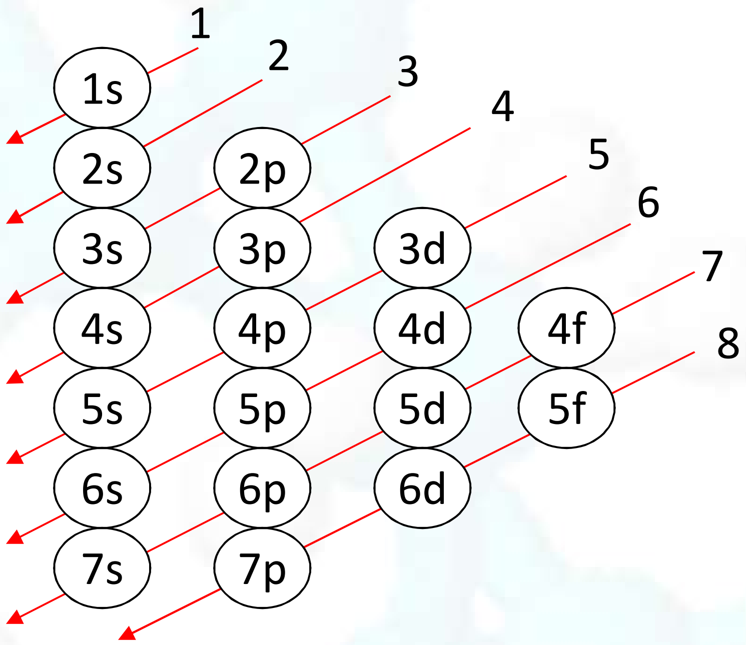
K atom has 19 electrons

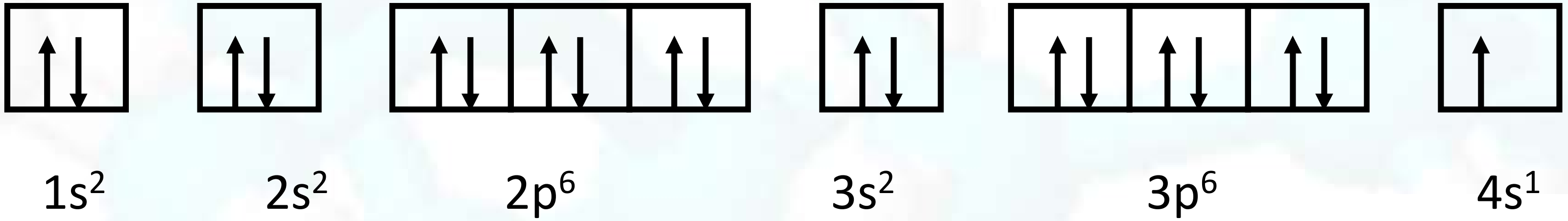
(from Periodic table (atomic number))

The Electron Configuration: $1s^2 2s^2 2p^6 3s^2\ 3p^6\ 4s^1$

2 + 2 + 6 + 2 +6+1= 19 electrons

Orbital Diagram:





Orbital diagram

What is the electron configuration of K^{+1}?

Atomic number19

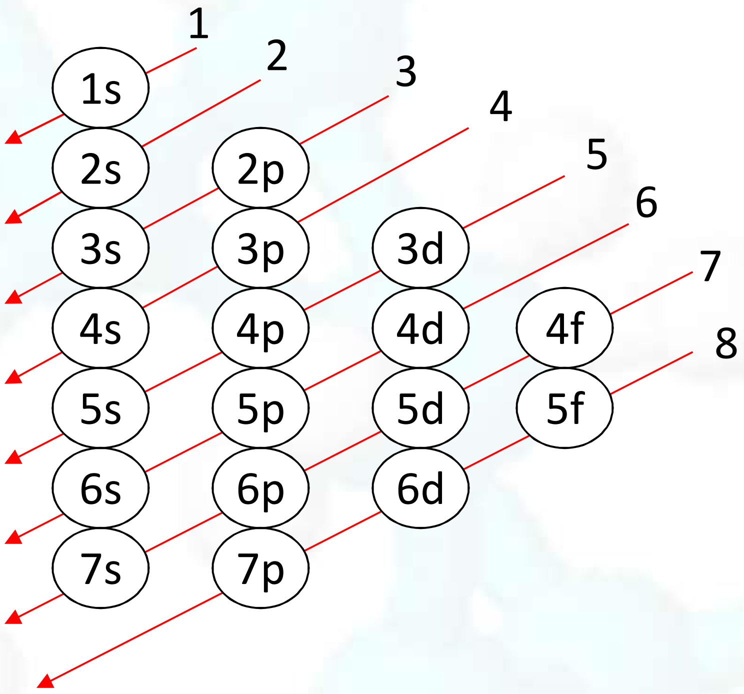
Electron 19 -1= 18

K atom has 18 electrons

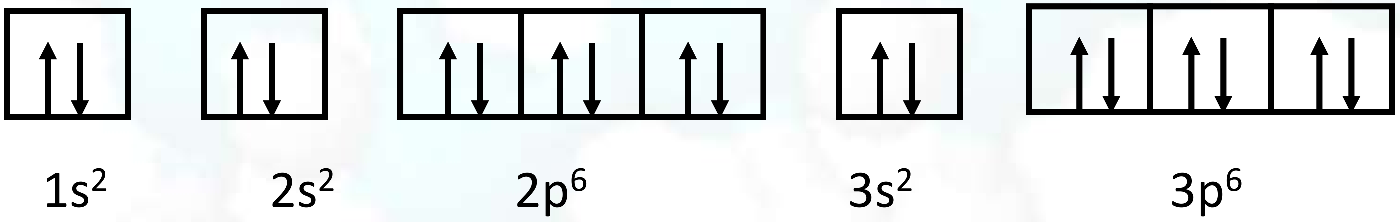
(from Periodic table (atomic number))

The Electron Configuration: $1s^2 2s^2 2p^6 3s^2\ 3p^6$

2 + 2 + 6 + 2 +6= 18 electrons

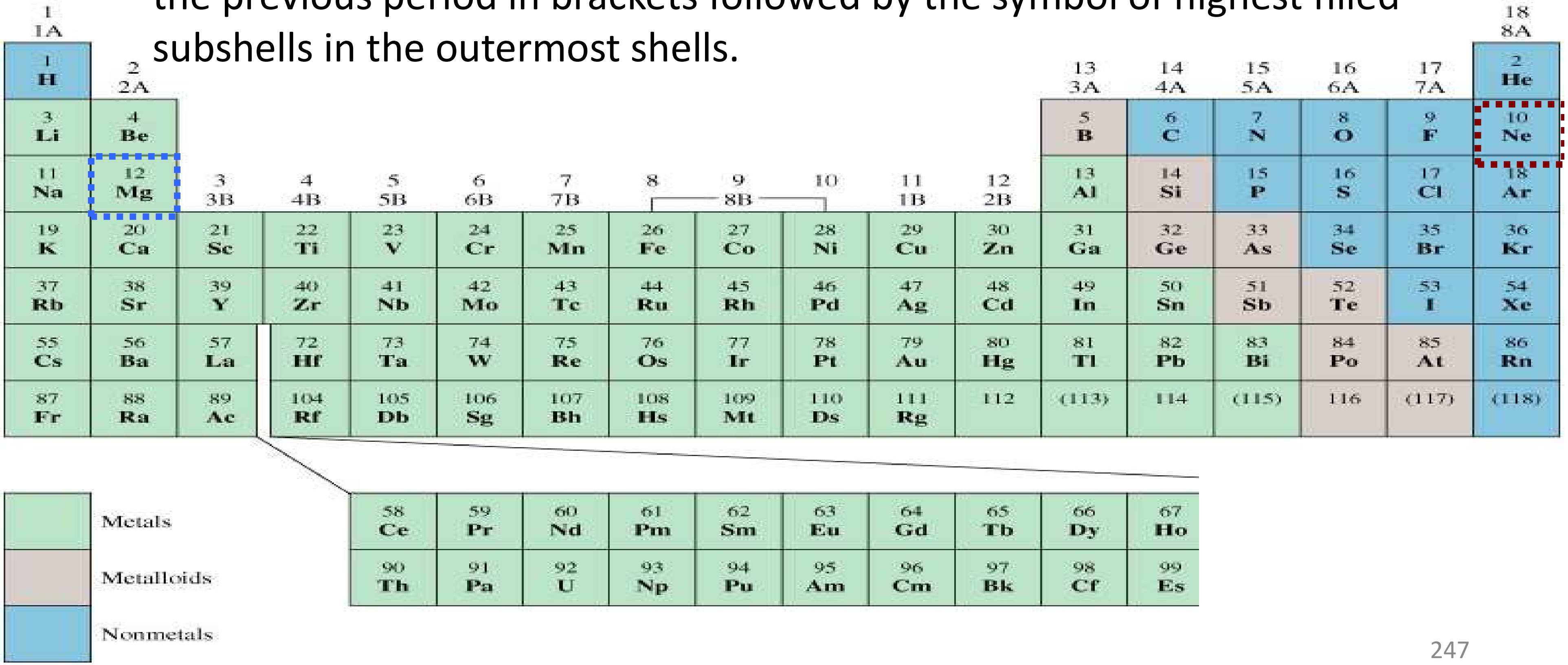


Orbital Diagram:



Short Notation

Short Notation (abbreviation): To write the electron configuration of an element in short notation, write the symbol of the Noble gas element in the previous period in brackets followed by the symbol of highest filled subshells in the outermost shells.



1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho				
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es				

🟩 Metals

🟫 Metalloids

🟦 Nonmetals

Atomic Orbitals

Example

For Cl atom answer the following questions:

a) Write the electron configuration?

b) Draw the orbital diagram?

c) Write the electron configuration in short notation?

d) What are the possible quantum numbers for the last (outermost) electron in Cl?

Answer:

a) Cl atom has 17 electrons

(from Periodic table (atomic number))

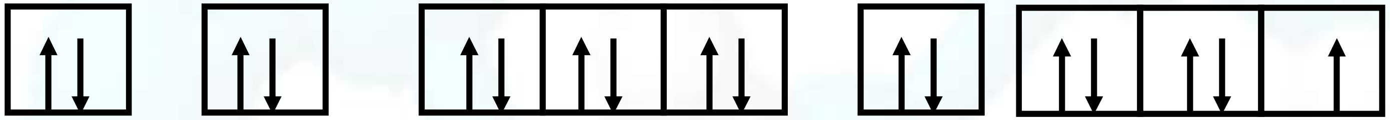
The Electron Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

$2 + 2 + 6 + 2\ \ + 5 = 17$ electrons

Atomic Orbitals

b) Draw the orbital diagram?

Orbital Diagram:



$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^5$

c) Short notation: [Ne] $3s^2$ $3p^5$

d) $n= 3$

$l= 1$

$m_l= 0$

$m_s= - ½$

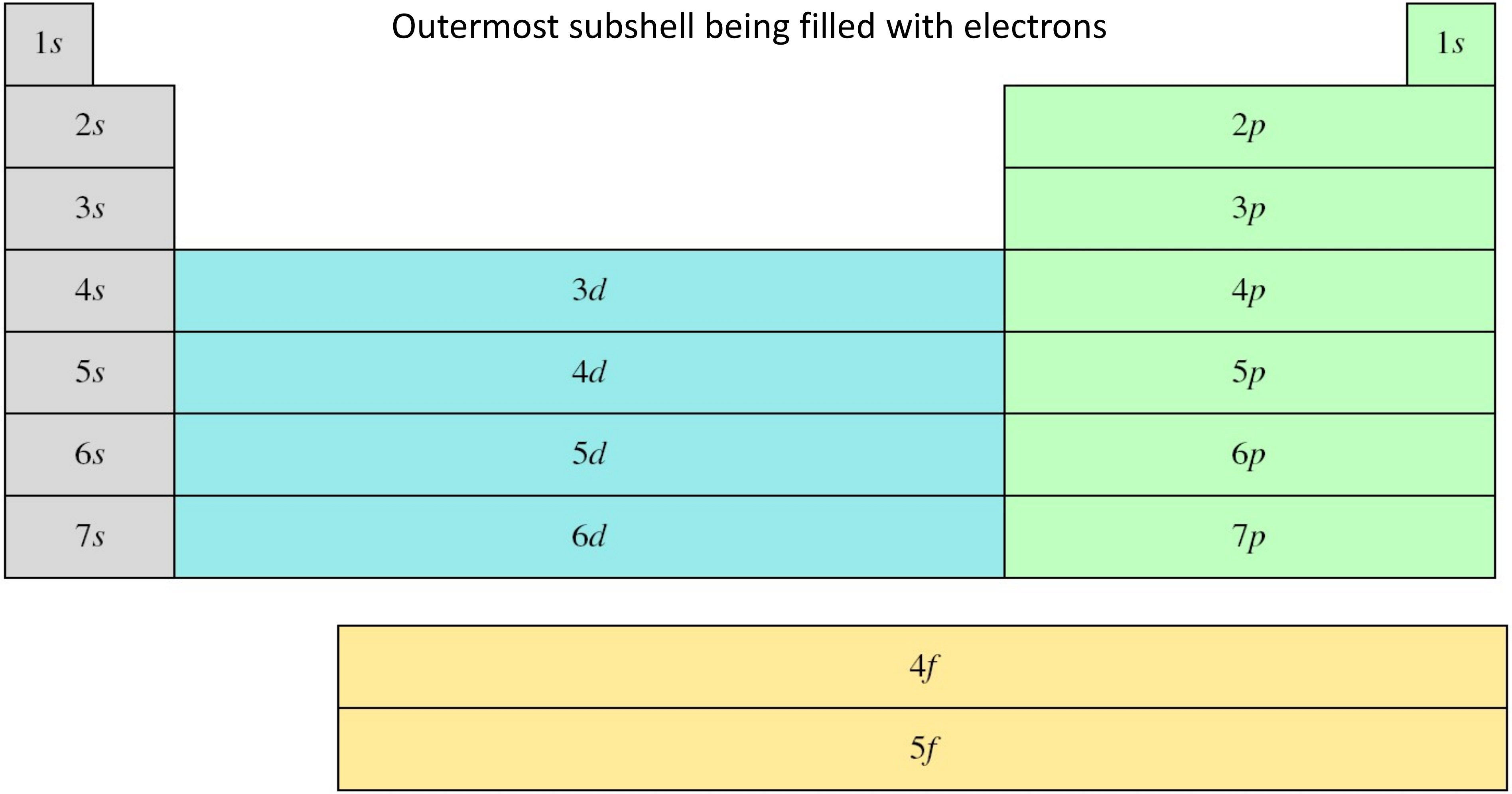
Atomic Orbitals

Example

What is The electron configuration of ^{11}Na , ^{12}Mg , ^{16}S ?

- ^{11}Na $1s^2\ 2s^2\ 2p^6\ 3s^1$ (OR) [Ne] $3s^1$
- ^{12}Mg $1s^2\ 2s^2\ 2p^6\ 3s^2$ (OR) [Ne] $3s^2$
- ^{16}S $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ (OR) [Ne] $3s^2\ 3p^4$

The Building-Up Principal



Outermost subshell being filled with electrons

1s		1s
2s		2p
3s		3p
4s	3d	4p
5s	4d	5p
6s	5d	6p
7s	6d	7p

4f
5f

Development of the Periodic Table

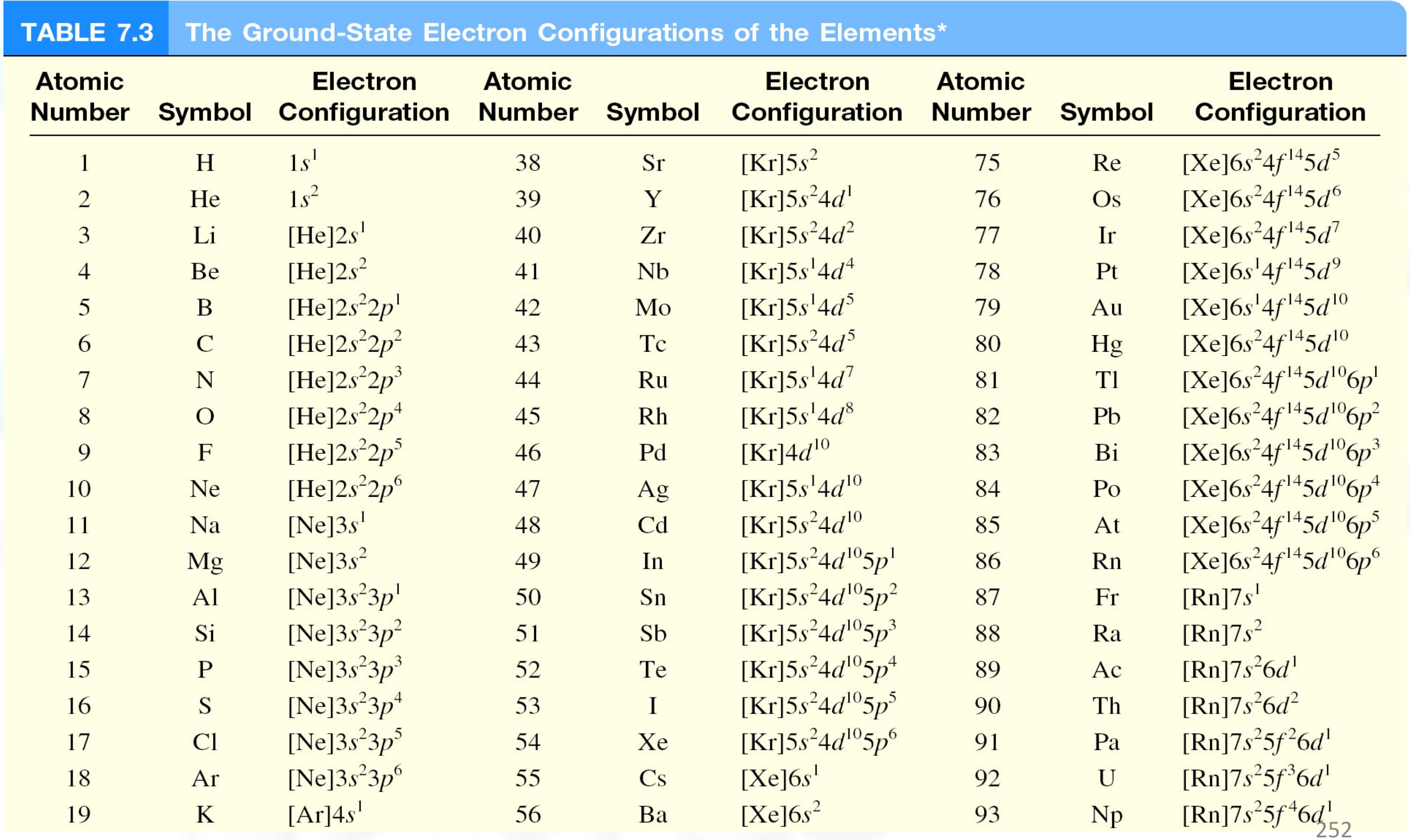
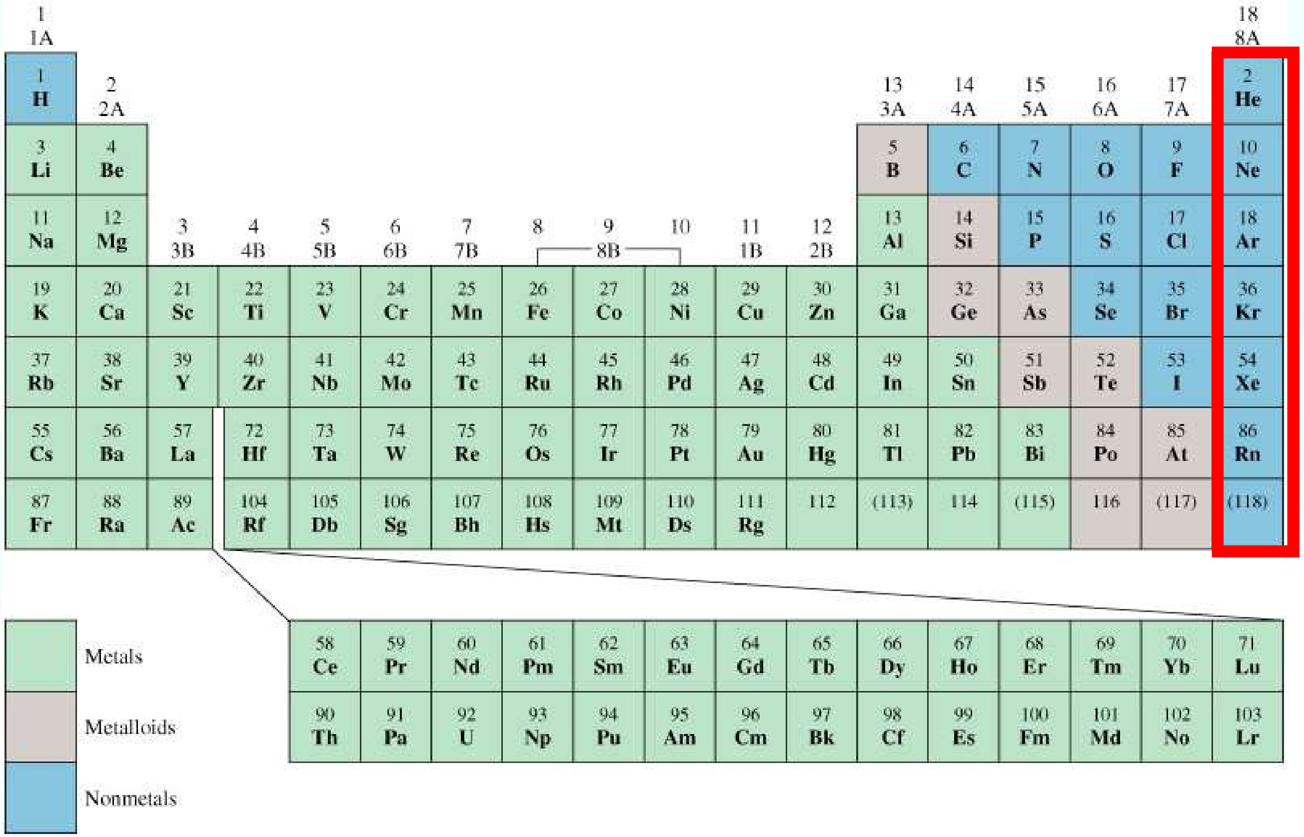


TABLE 7.3 The Ground-State Electron Configurations of the Elements*

Atomic Number	Symbol	Electron Configuration	Atomic Number	Symbol	Electron Configuration	Atomic Number	Symbol	Electron Configuration
1	H	$1s^1$	38	Sr	$[Kr]5s^2$	75	Re	$[Xe]6s^24f^{14}5d^5$
2	He	$1s^2$	39	Y	$[Kr]5s^24d^1$	76	Os	$[Xe]6s^24f^{14}5d^6$
3	Li	$[He]2s^1$	40	Zr	$[Kr]5s^24d^2$	77	Ir	$[Xe]6s^24f^{14}5d^7$
4	Be	$[He]2s^2$	41	Nb	$[Kr]5s^14d^4$	78	Pt	$[Xe]6s^14f^{14}5d^9$
5	B	$[He]2s^22p^1$	42	Mo	$[Kr]5s^14d^5$	79	Au	$[Xe]6s^14f^{14}5d^{10}$
6	C	$[He]2s^22p^2$	43	Tc	$[Kr]5s^24d^5$	80	Hg	$[Xe]6s^24f^{14}5d^{10}$
7	N	$[He]2s^22p^3$	44	Ru	$[Kr]5s^14d^7$	81	Tl	$[Xe]6s^24f^{14}5d^{10}6p^1$
8	O	$[He]2s^22p^4$	45	Rh	$[Kr]5s^14d^8$	82	Pb	$[Xe]6s^24f^{14}5d^{10}6p^2$
9	F	$[He]2s^22p^5$	46	Pd	$[Kr]4d^{10}$	83	Bi	$[Xe]6s^24f^{14}5d^{10}6p^3$
10	Ne	$[He]2s^22p^6$	47	Ag	$[Kr]5s^14d^{10}$	84	Po	$[Xe]6s^24f^{14}5d^{10}6p^4$
11	Na	$[Ne]3s^1$	48	Cd	$[Kr]5s^24d^{10}$	85	At	$[Xe]6s^24f^{14}5d^{10}6p^5$
12	Mg	$[Ne]3s^2$	49	In	$[Kr]5s^24d^{10}5p^1$	86	Rn	$[Xe]6s^24f^{14}5d^{10}6p^6$
13	Al	$[Ne]3s^23p^1$	50	Sn	$[Kr]5s^24d^{10}5p^2$	87	Fr	$[Rn]7s^1$
14	Si	$[Ne]3s^23p^2$	51	Sb	$[Kr]5s^24d^{10}5p^3$	88	Ra	$[Rn]7s^2$
15	P	$[Ne]3s^23p^3$	52	Te	$[Kr]5s^24d^{10}5p^4$	89	Ac	$[Rn]7s^26d^1$
16	S	$[Ne]3s^23p^4$	53	I	$[Kr]5s^24d^{10}5p^5$	90	Th	$[Rn]7s^26d^2$
17	Cl	$[Ne]3s^23p^5$	54	Xe	$[Kr]5s^24d^{10}5p^6$	91	Pa	$[Rn]7s^25f^26d^1$
18	Ar	$[Ne]3s^23p^6$	55	Cs	$[Xe]6s^1$	92	U	$[Rn]7s^25f^36d^1$
19	K	$[Ar]4s^1$	56	Ba	$[Xe]6s^2$	93	Np	$[Rn]7s^25f^46d^1$

Atomic Orbitals

- Noble Gases: elements with electron configuration of complete s & p subshell (He, Ne, Ar, Kr, Xe, Rn), Group 8A

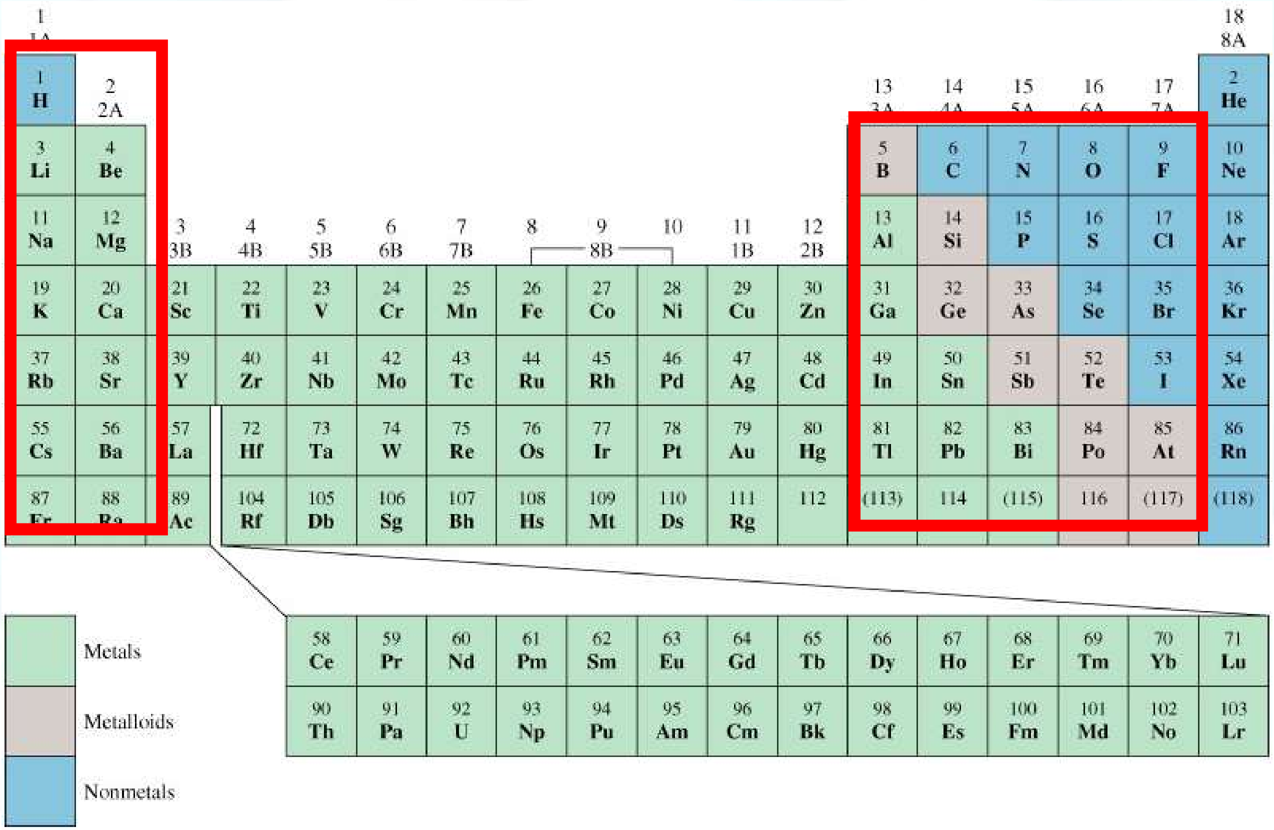


1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H**																	2 **He**
3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
55 **Cs**	56 **Ba**	57 **La**	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
87 **Fr**	88 **Ra**	89 **Ac**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112	(113)	114	(115)	116	(117)	(118)
				58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
				90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

🟩 Metals
⬜ Metalloids
🟦 Nonmetals

Atomic Orbitals

- Representative Elements: elements with electron configuration of incompletely filled s or p subshell. Elements in Groups 1A-to-7A



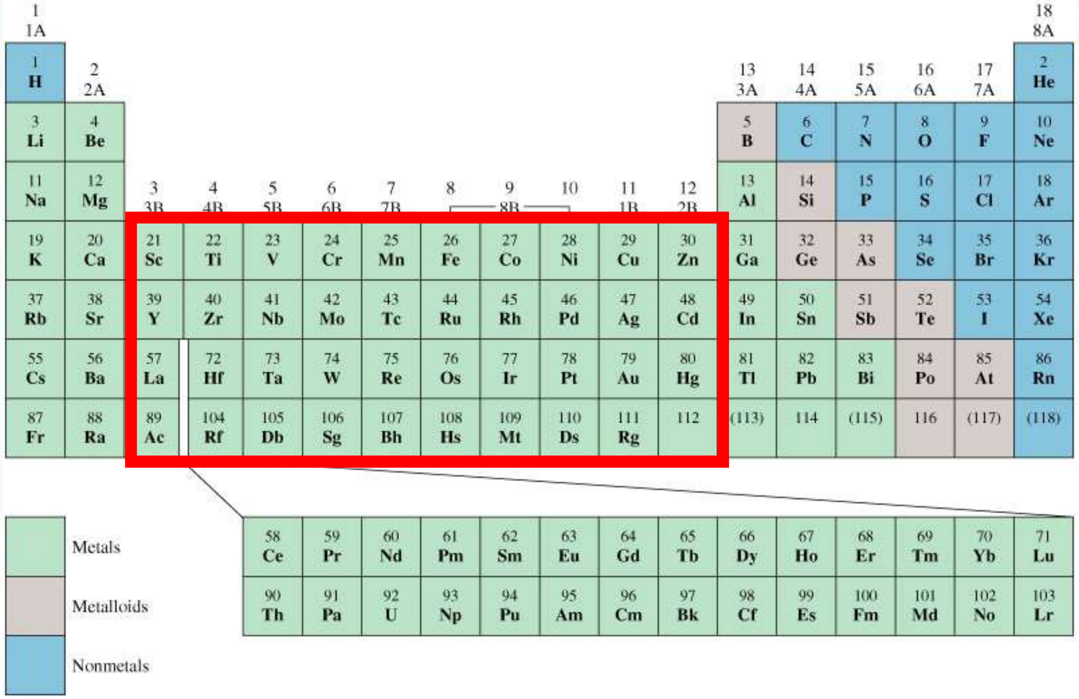
1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals
▨ Metalloids
🟦 Nonmetals

Atomic Orbitals

- Transition Elements: elements with electron configuration of incompletely filled d subshells or readily give rise to cations that have incompletely filled d subshells. Elements in Groups 1B-to-7B



1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals
▨ Metalloids
🟦 Nonmetals

Exceptions

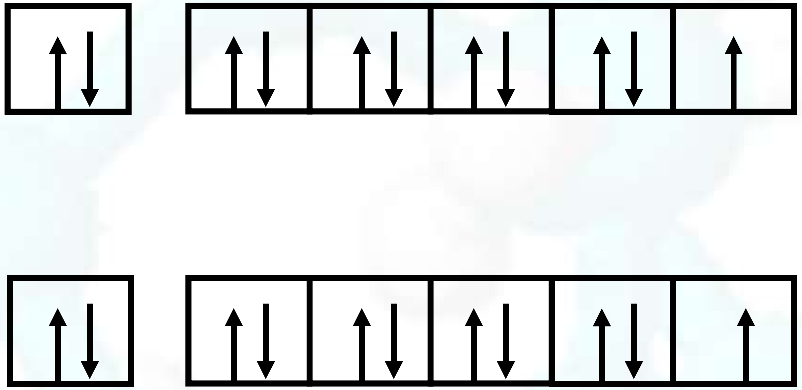
The Stability of Half Filled & Filled d Orbitals

$^{29}\text{Cu}:1\underline{\text{s2 2s2 2p6 3s2 3p6}}\ 4s^2\ 3d^9$

$[\text{Ar}]\ 4s^2\ 3d^9$

The stability of Filled d orbital →

$^{29}\text{Cu}:\ [\text{Ar}]\ 4s^1\ 3d^{10}$



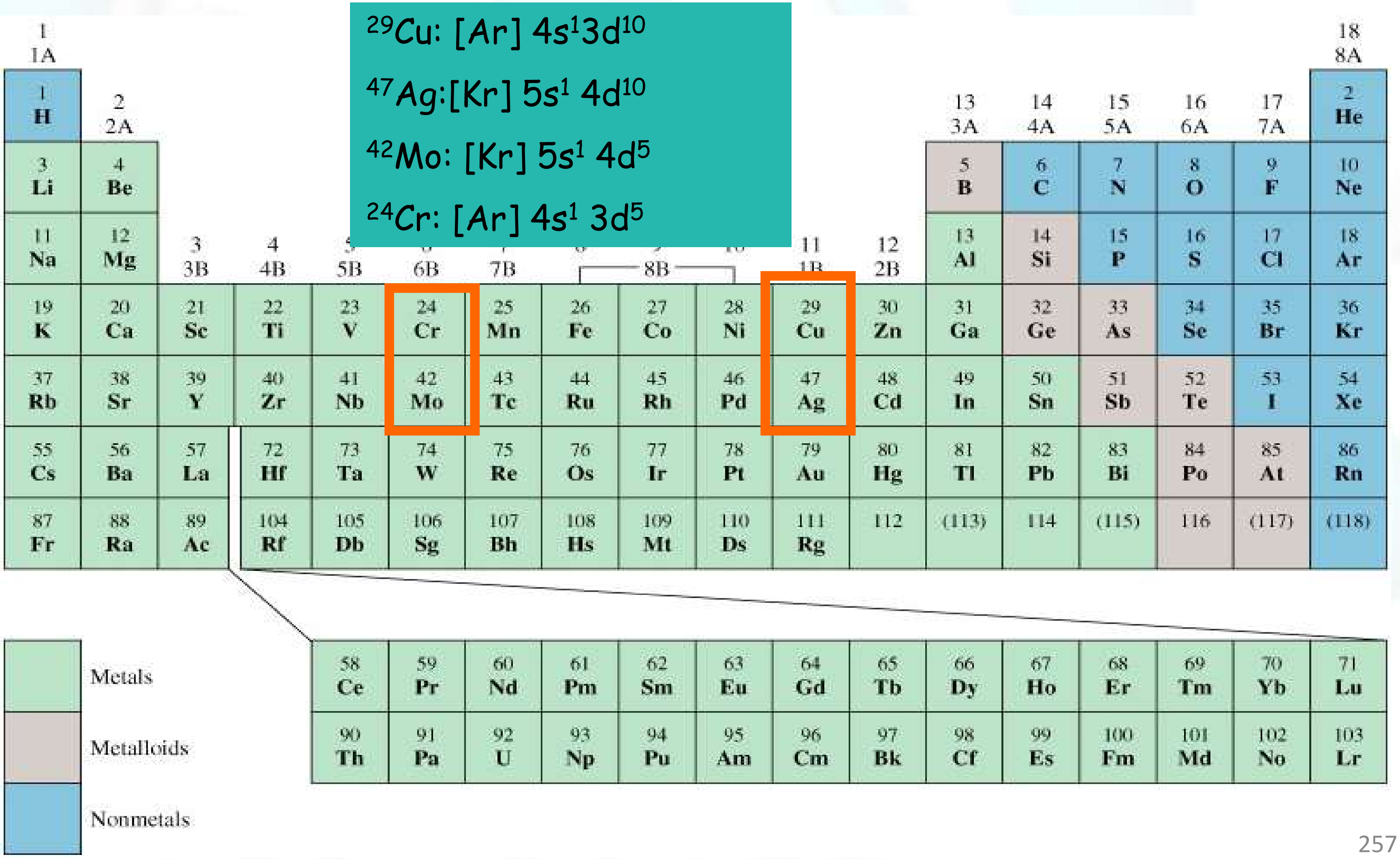
Atomic Orbitals

^{29}Cu: [Ar] $4s^{1}3d^{10}$

^{47}Ag:[Kr] $5s^{1}$ $4d^{10}$

^{42}Mo: [Kr] $5s^{1}$ $4d^{5}$

^{24}Cr: [Ar] $4s^{1}$ $3d^{5}$



1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B		8B		11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)
			58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
			90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

🟩 Metals

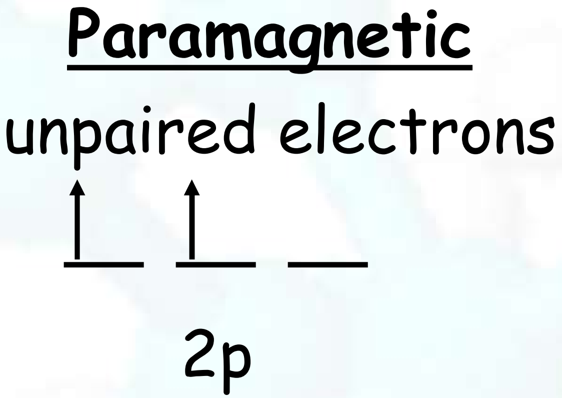
▨ Metalloids

🟦 Nonmetals

Atomic Orbitals

- Paramagnetic substance: is the element that contain net unpaired electrons in the outermost subshell and is attracted by a magnet.

e.g.



Paramagnetic

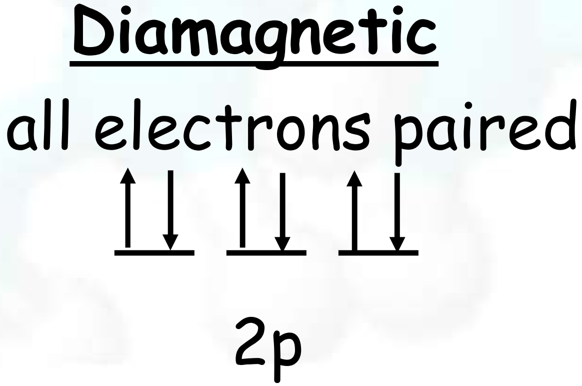
unpaired electrons

↑ ↑ __

2p

- Diamagnetic substance: is the element that do not contain net unpaired electrons (all electrons are paired) in the outermost subshell and is repelled by a magnet.

e.g.



Diamagnetic

all electrons paired

↑↓ ↑↓ ↑↓

2p

Atomic Orbitals

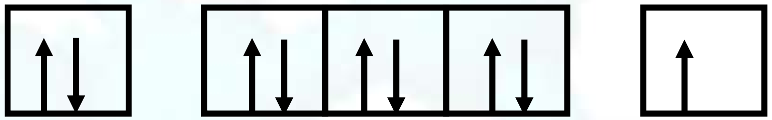
^{19}K

Example

Electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

Short notation: $[Ar]\ 4s^1$

Orbital diagram:



Net one unpaired electron → Paramagnetic substance

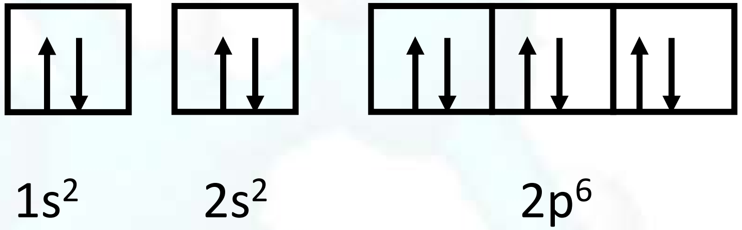
Atomic Orbitals

Example

^{10}Ne

Electronic configuration: $1s^2\ 2s^2\ 2p^6$

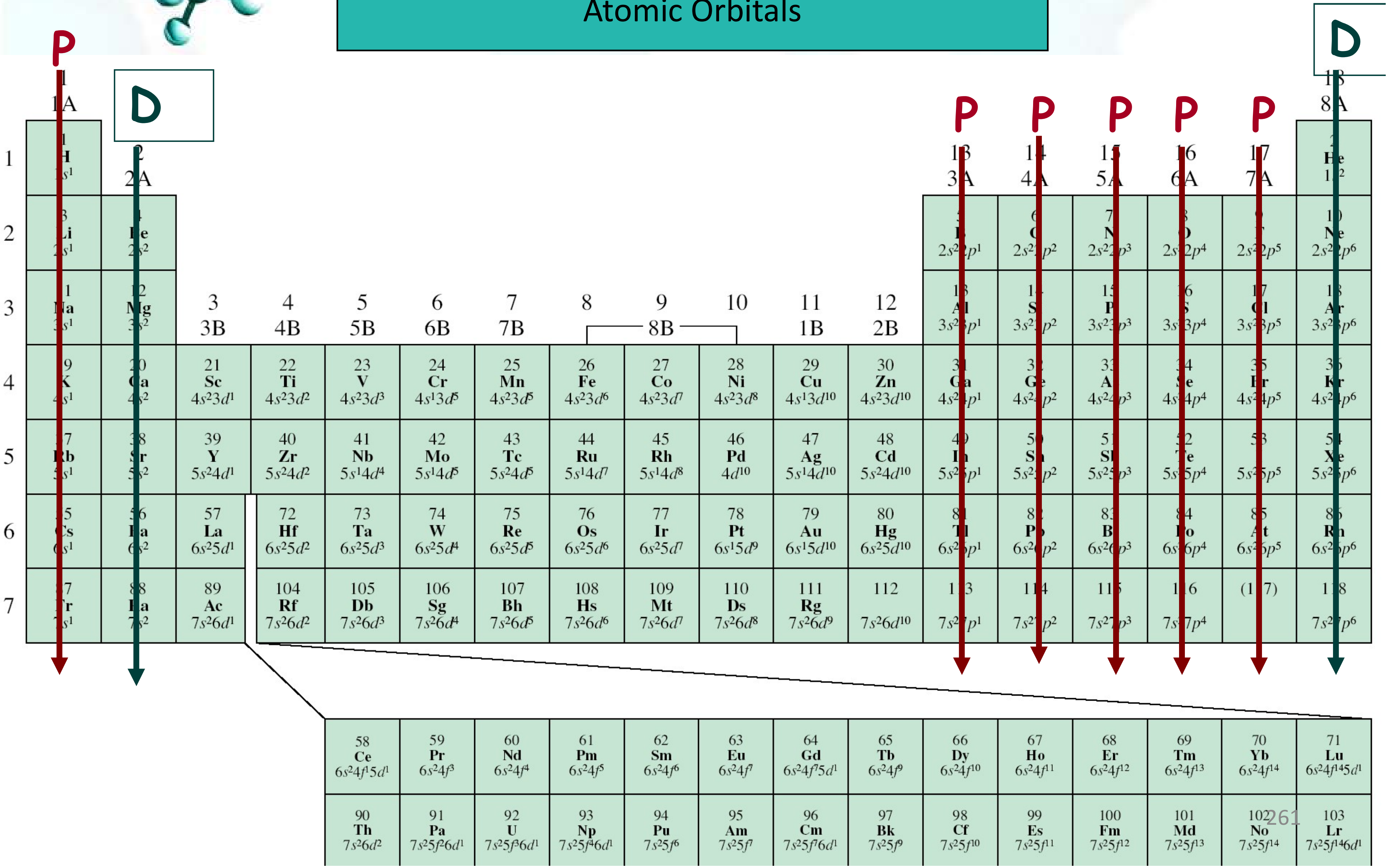
Orbital diagram:



All electrons are paired →

Diamagnetic substance

Atomic Orbitals



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1	1 **H** $1s^1$																	2 **He** $1s^2$
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
3	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh** $7s^26d^5$	108 **Hs** $7s^26d^6$	109 **Mt** $7s^26d^7$	110 **Ds** $7s^26d^8$	111 **Rg** $7s^26d^9$	112 $7s^26d^{10}$	113 $7s^27p^1$	114 $7s^27p^2$	115 $7s^27p^3$	116 $7s^27p^4$	(117)	118 $7s^27p^6$

58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

Atomic Orbitals

Example

What are the valence electrons of vanadium (V)?

^{23}V: [Ar] $4s^2$ $3d^3$

Example

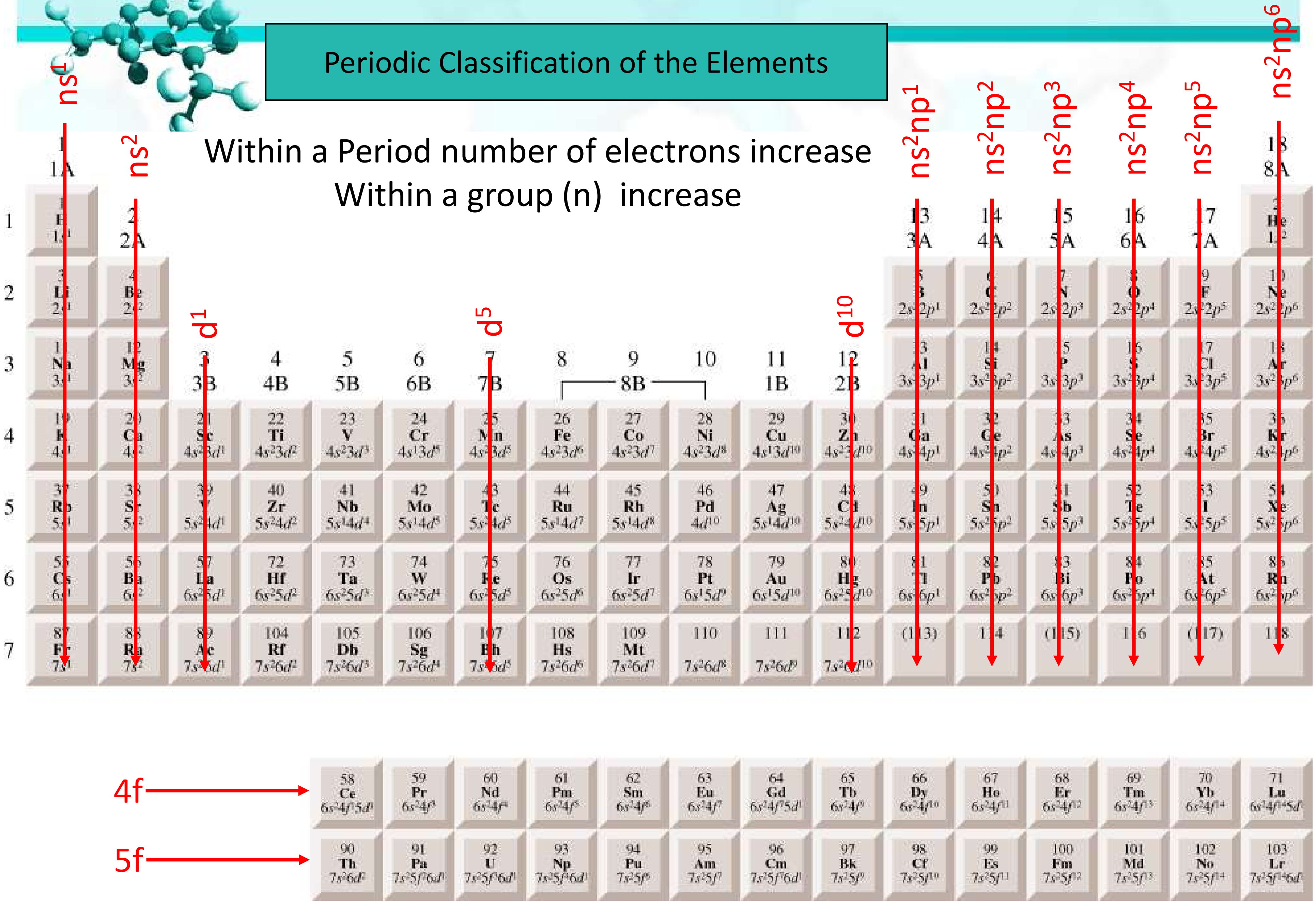
What are the valence electrons of Gallium (Ga)?

^{31}Ga: [Ar] $4s^2$ $3d^{10}$ $4p^1$

The valence electrons are $4s^2$ $4p^1$

Periodic Classification of the Elements

Within a Period number of electrons increase
Within a group (n) increase



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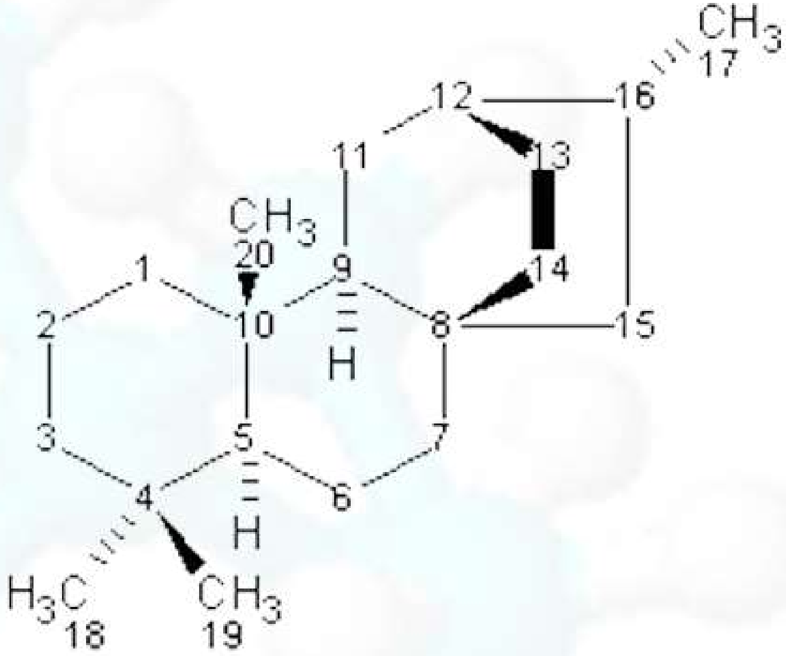
1	1 **H** $1s^1$																	2 **He** $1s^2$
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
3	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh** $7s^26d^5$	108 **Hs** $7s^26d^6$	109 **Mt** $7s^26d^7$	110 $7s^26d^8$	111 $7s^26d^9$	112 $7s^26d^{10}$	(113)	114	(115)	116	(117)	118
4f					58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
5f					90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

Thank you

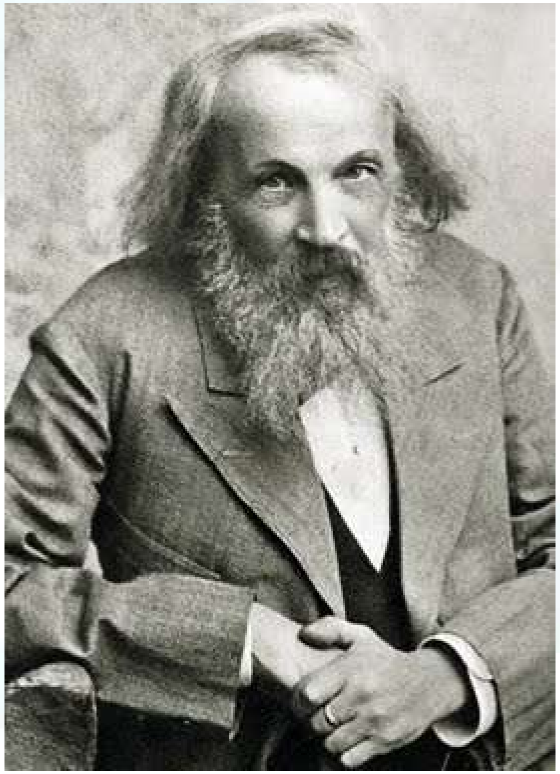


Chapter Eight

Periodic Relationships Among the Elements

Development of the Periodic Table

- Many attempt has been made to arrange the element.
- Russian Chemist Dmitri Mendeleev arrange the element based on the regular, periodic recurrence of properties.
- Mendeleev’s classification system was a great improvement for two reasons. First, it grouped the elements together more accurately, according to their properties. Equally important, it made possible the prediction of the properties of several elements that had not yet been discovered.



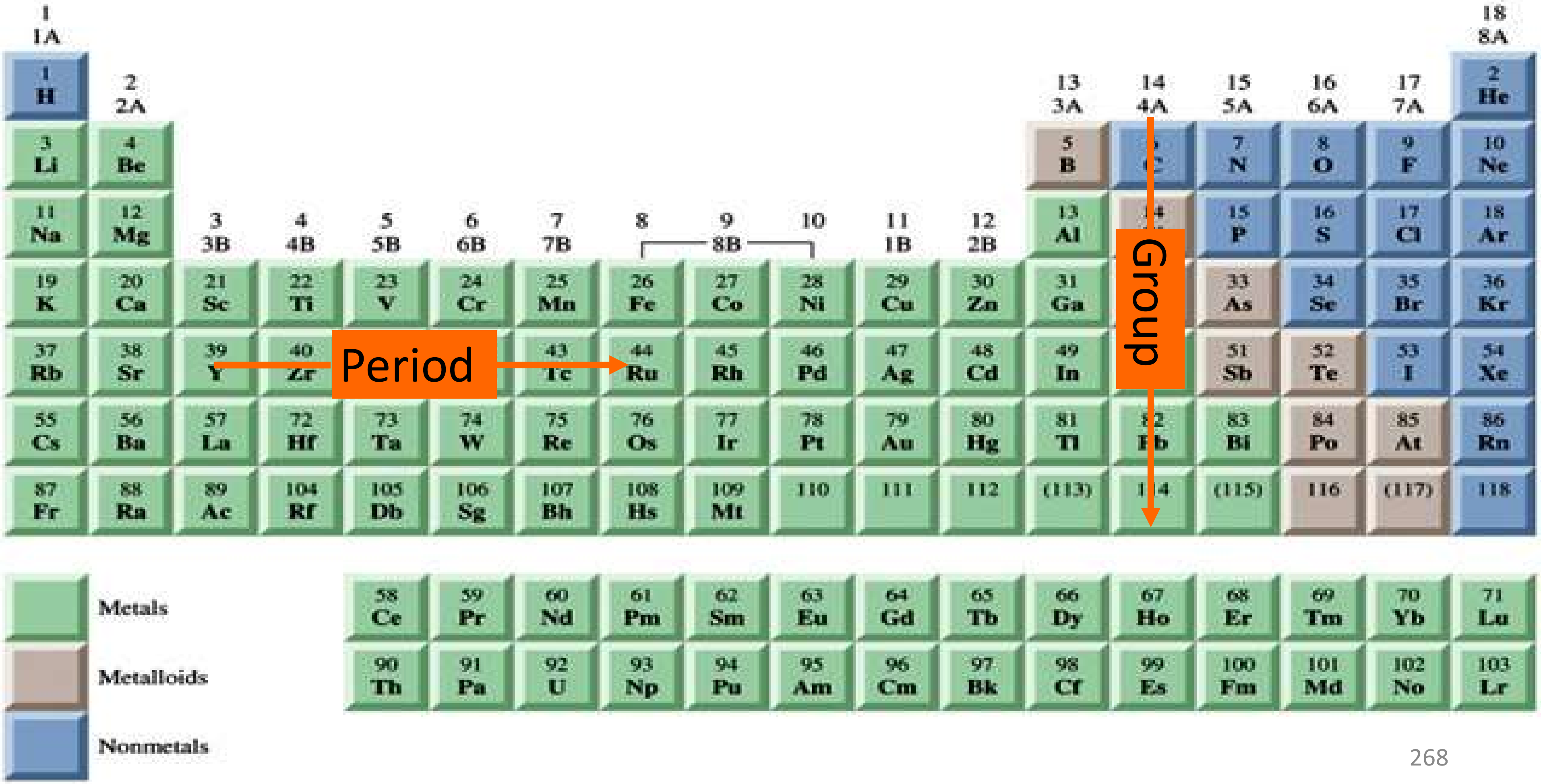
Dmitri Mendeleev
1834-1907

Development of the Periodic Table

- Mendeleev’s classification was based on the atomic weight, however, this resulted in some inconsistency.
- Later on after the discovery of atomic number by Henry Moseley the element were arranged by their atomic number.

Periodic Classification of the Elements

Modern Periodic Table



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1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	[illegible]	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	[illegible]	[illegible]	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	[illegible]	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112	(113)	114	(115)	116	(117)	118
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals

🟫 Metalloids

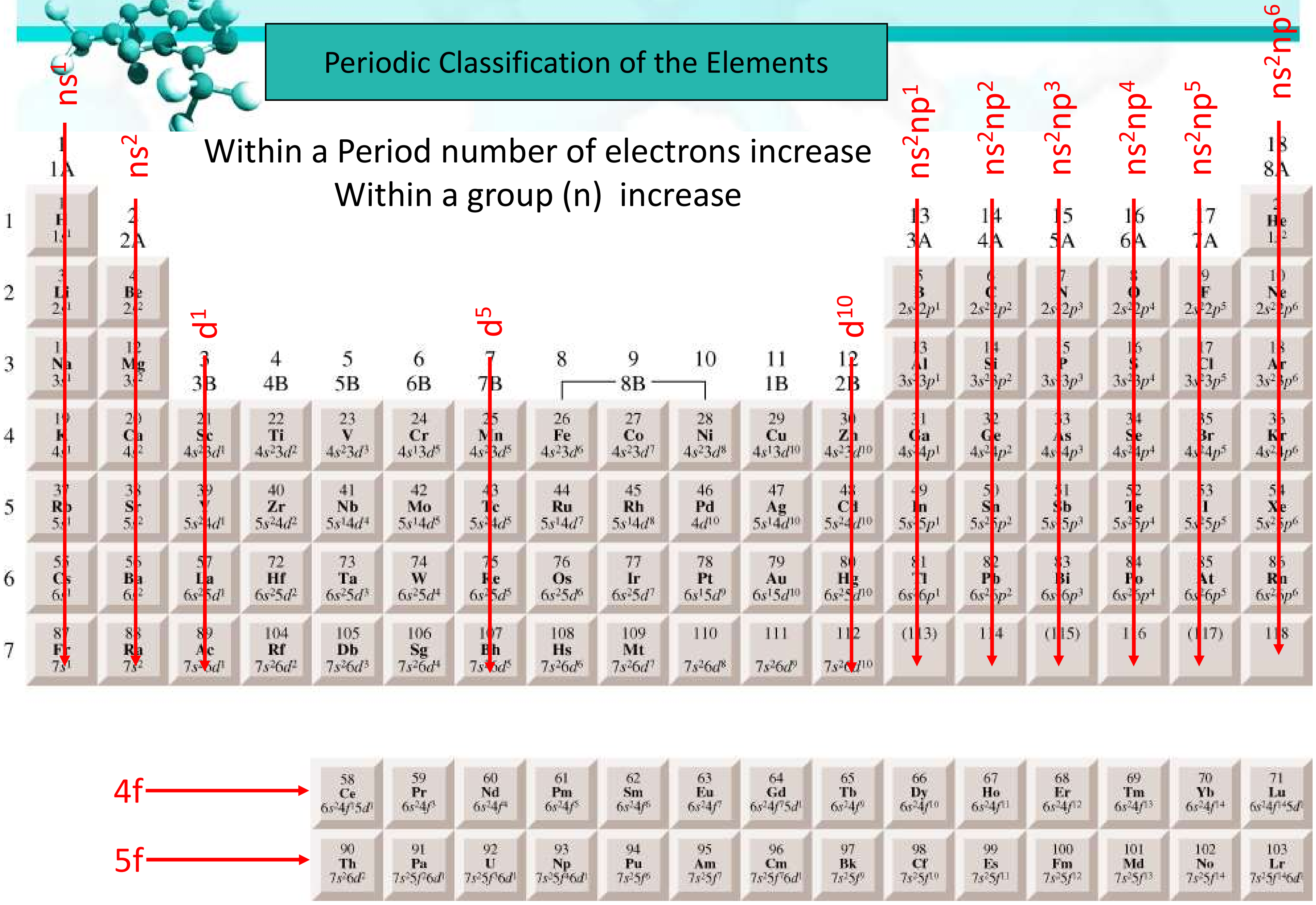
🟦 Nonmetals

Period

Group

Periodic Classification of the Elements

Within a Period number of electrons increase
Within a group (n) increase



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1	1 **H** $1s^1$																	2 **He** $1s^2$
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
3	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh** $7s^26d^5$	108 **Hs** $7s^26d^6$	109 **Mt** $7s^26d^7$	110 $7s^26d^8$	111 $7s^26d^9$	112 $7s^26d^{10}$	(113)	114	(115)	116	(117)	118

58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
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Periodic Classification of the Elements

Example:

The element that has the valance electron configuration $3s^2\ 3p^3$ is:

a)Carbon b) Nitrogen [✓ c) Phosphorus] d)Neon

Periodic Classification of the Elements

Example 2:

An atom of a certain element has 15 electrons. Without consulting a periodic table, answer the following questions:

(a) What is the ground-state electron configuration of this element?

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

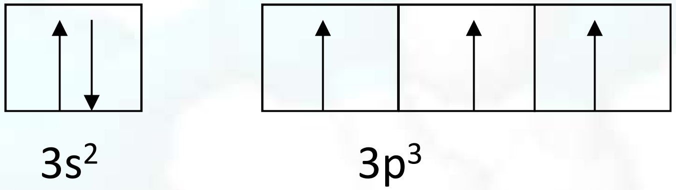
(b) How should be element be classified?

Period 3, group 5A

The element is representative element.

(c) Is the element diamagnetic or paramagnetic

paramagnetic

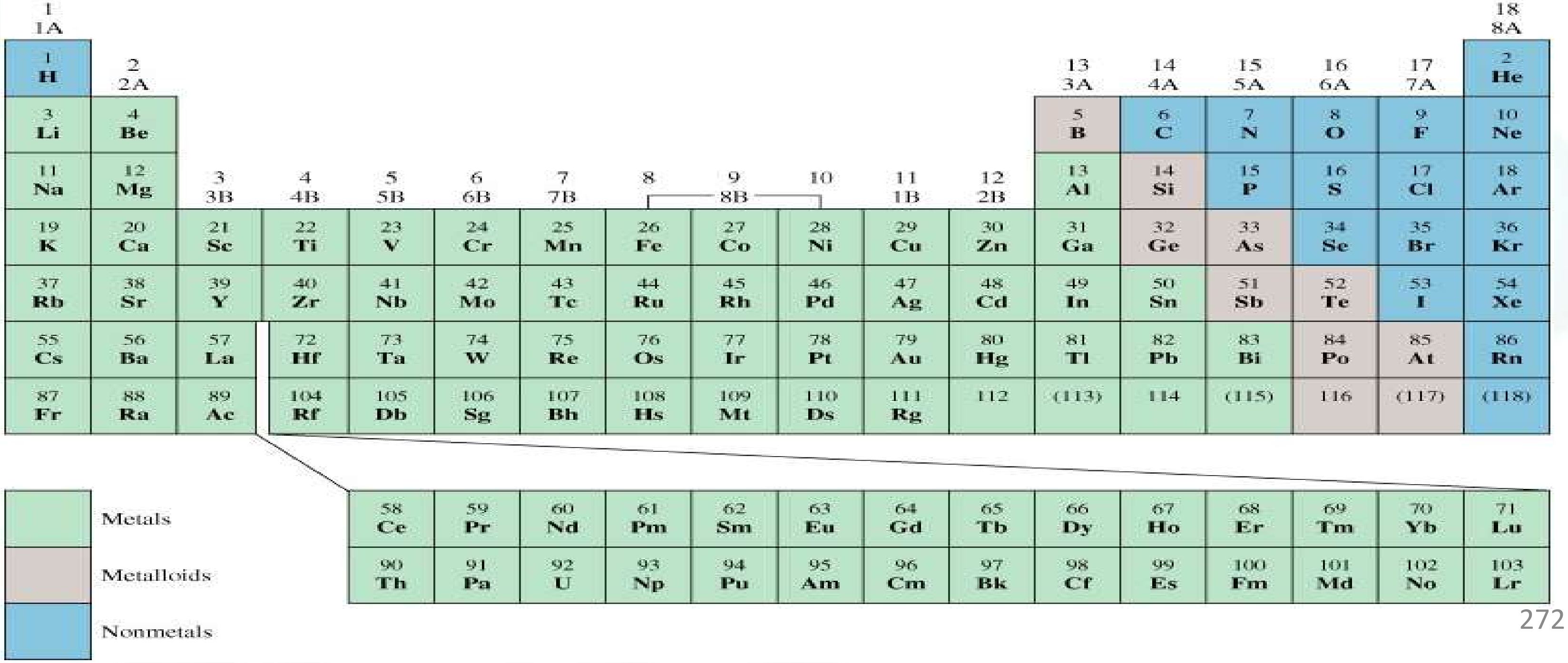


Development of the periodic table

Example:

Which of the following sets of elements is expected to have similar chemical properties?

a) Sulfur and phosphorous b) Sulfur and oxygen c) Sulfur and argon



1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
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58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals

🟫 Metalloids

🟦 Nonmetals

Periodic Classification of the Elements

Example

Titanium (Ti) element is found in the periodic table in

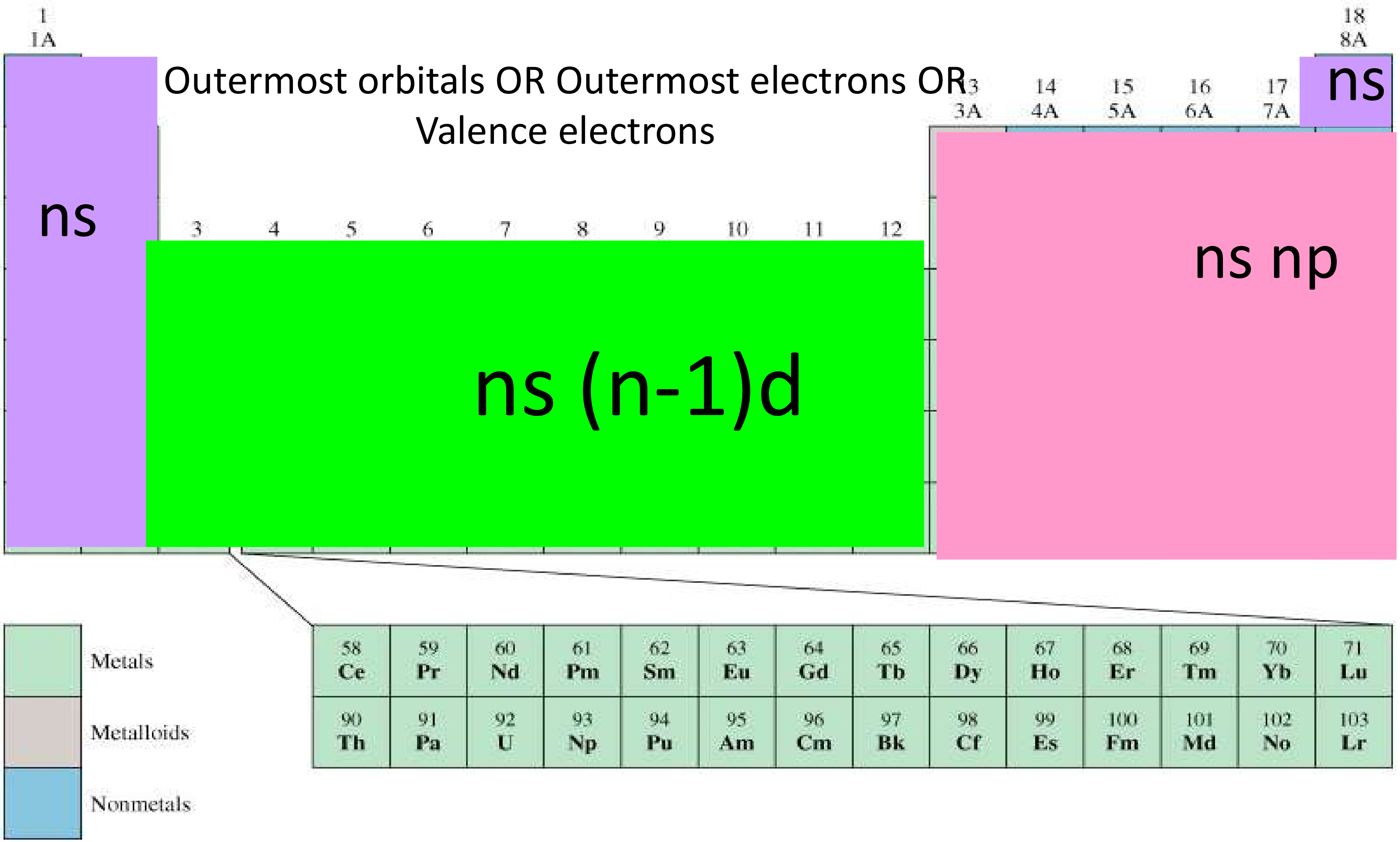
s-block (b) P-block (c) [✓ d-block] (d) f-block

Example

Characteristics of noble gases include:

- a. filled s and p subshells.
- b. monatomic gases.
- c. generally unreactive chemical.
- [✓ d. all of the above.]

Periodic Classification of the Elements



Outermost orbitals OR Outermost electrons OR Valence electrons

Electron Configration of Cation and Anio

Ion derived from representative element

In the formation of a cation from the atom of a representative element, one or more electrons are removed from the highest occupied n shell so that Cation has a noble-gas outer electron configuration. The electron configurations of some atoms and their corresponding cations are as follows:

Na: [Ne] $3s^1$ Na$^+$ [Ne]

Ca: [Ar] $4s^2$ Ca^{2+} [Ar]

Al: [Ne] $3s^2 3p^1$ Al^{3+} [Ne]

Electron Configration of Cation and Anio

Ion derived from representative element

In the formation of an anion, one or more electrons are added to the highest Partially filled n shell so that anion has a noble-gas outer electron configuration. Consider the following examples:

H: $1s^1$	H^- $1s^2$ or [He]
F: $1s^22s^22p^5$	F^- $1s^22s^22p^6$ or [Ne]
O: $1s^22s^22p^4$	O^{2-} $1s^22s^22p^6$ or [Ne]
N: $1s^22s^22p^3$	N^{3-} $1s^22s^22p^6$ or [Ne]

Electron Configration of Cation and Anio

Ion derived from representative element

Isoelectronic: Species with the same number of electrons.

Example :

H^{-} :[He] , F^{-} : [Ne], N^{-3}: [Ne] , Na^{+} : [Ne] , Al^{+3} : [Ne], O^{-2} : [Ne]

^{11}Na (11 e) $\longrightarrow$ Na^{+} (10 e)

^{13}Al (13 e) $\longrightarrow$ Al^{3+} (10 e)

^{7}N (7 e) $\longrightarrow$ N^{3-} (10 e)

^{9}F (9 e) $\longrightarrow$ F^{-} (10 e)

^{8}O (8 e) $\longrightarrow$ O^{2-} (10 e)

^{10}Ne (10 e)

THUS: All are isolelctronic to Ne

Electron Configration of Cation and Anio

Ion derived from representative element

Example:

Which of the following species is isoelectronic with Cl^- 18 e

(a) F^- 10 e

(b) O^{2-} 10 e

$\boxed{\text{(c) } K^+ \quad 18\text{ e}}$

(d) Na^+ 10 e

Electron Configration of Cation and Anio

Ion derived from transition element

When a cation is formed from an atom of a transition metal, electrons are always removed first from the ns orbital and then from the (n –1)d orbitals.

Example:

Mn: $[Ar]4s^2 3d^5$

Mn^{2+}: $[Ar]\ 3d^5$

Fe: $[Ar]4s^2 3d^6$

Fe^{2+}: $[Ar]3d^6$

Fe^{3+}: $[Ar]3d^5$

Variation in physical prosperities

Effective nuclear charge

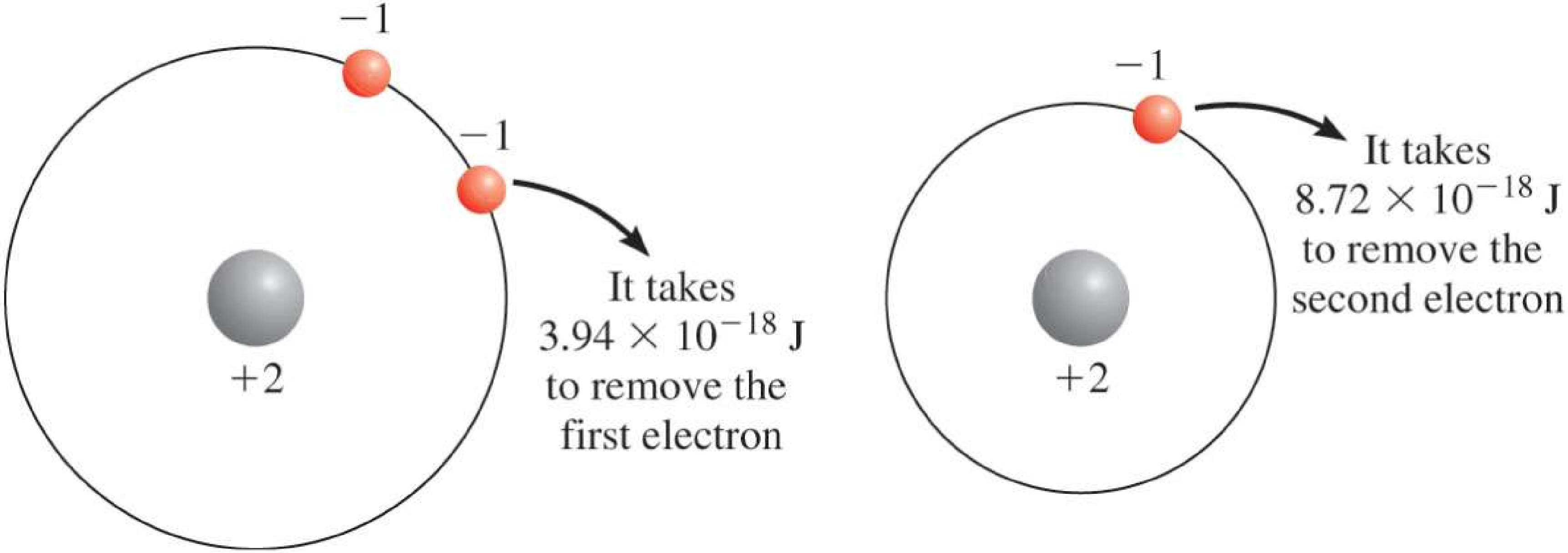
Effective nuclear charge(Z_{eff}): the nuclear charge felt by an electron when both the actual nuclear charge (Z) and the repulsive effects (shielding) of the other electrons are taken into account.

$$Z_{eff} = Z - \sigma$$

Where σ (sigma) is called the shielding constant.

Variation in physical prosperities

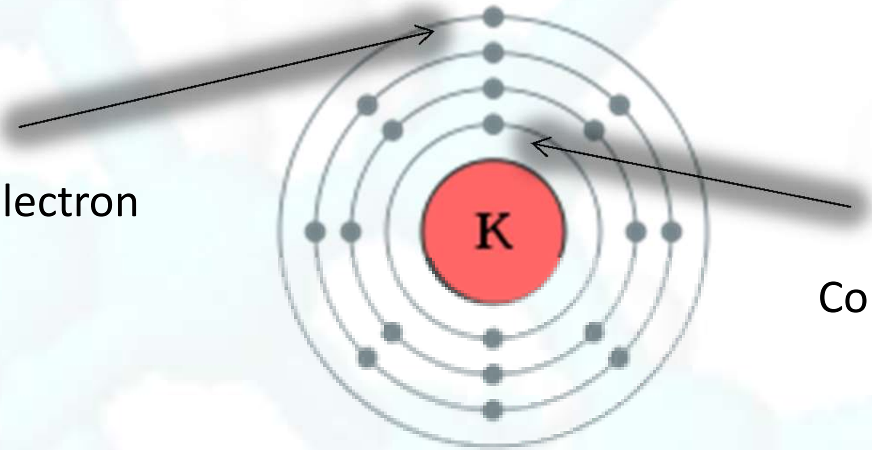
Effective nuclear charge



Variation in physical prosperities

Effective nuclear charge

19: Potassium **2,8,8,1**



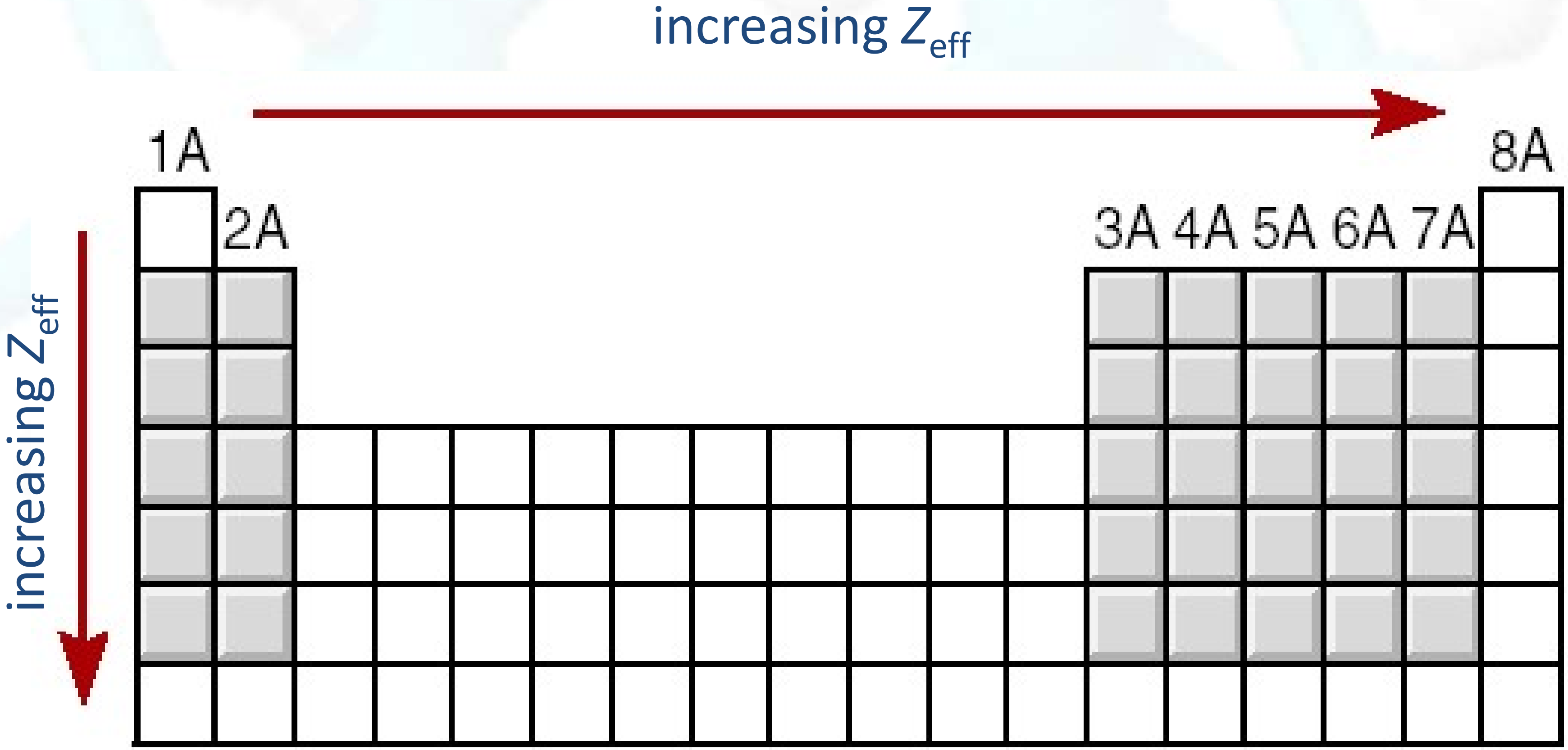
Valence electron

Core electron

The core electrons shield valence electrons MUCH MORE than valence electrons shield one another.

Variation in physical prosperities

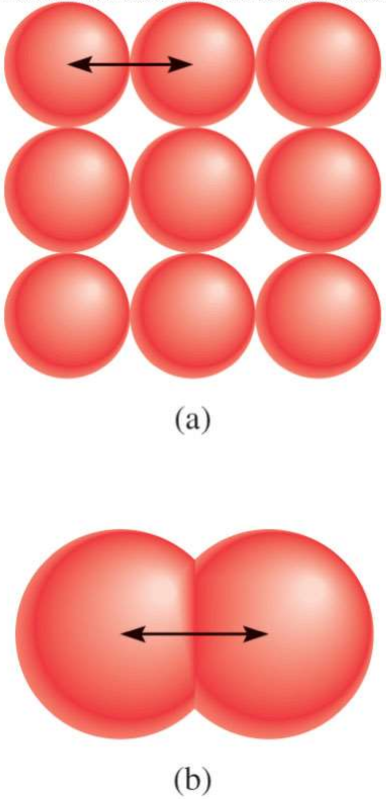
Effective nuclear charge



Variation in physical prosperities

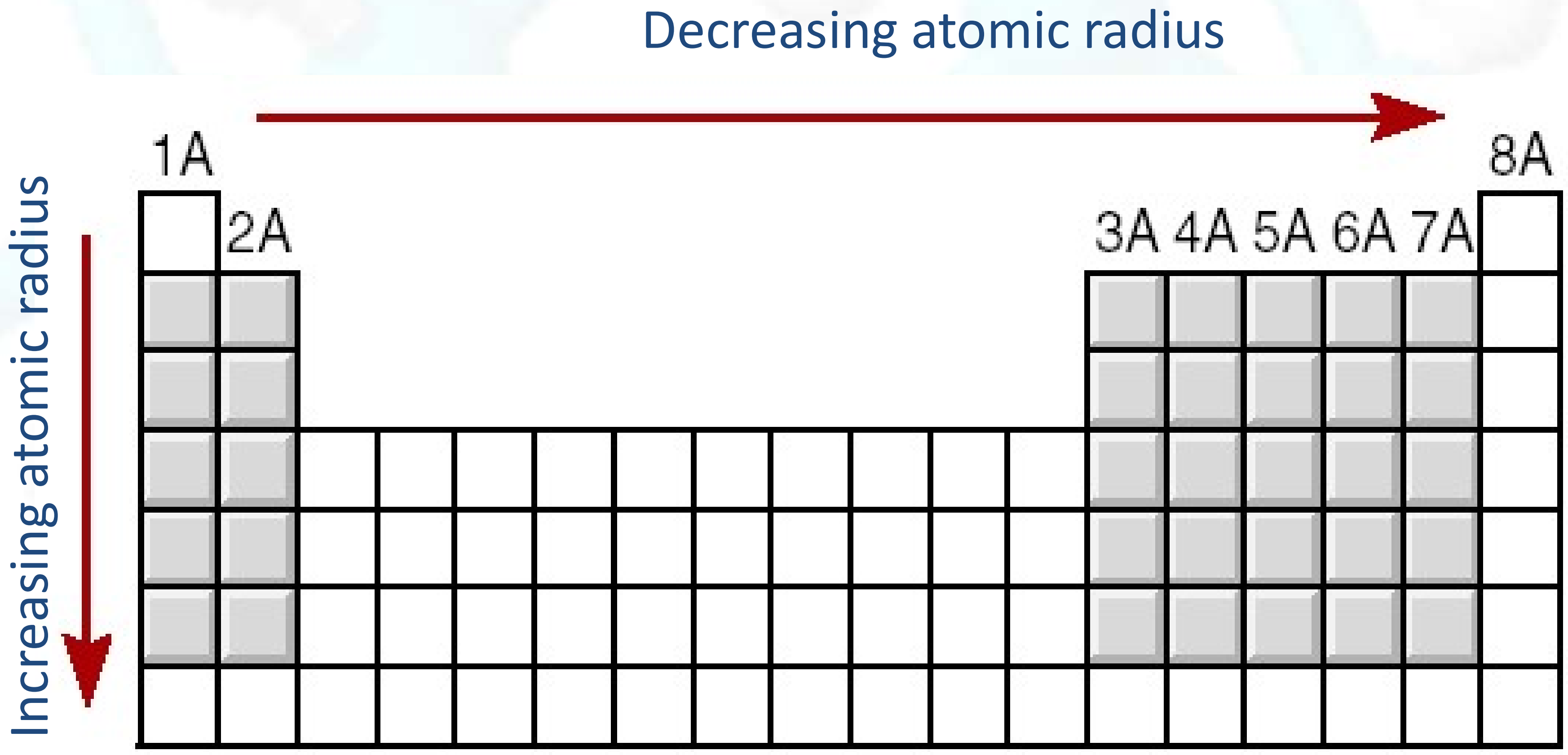
Atomic Radius

Atomic Radius: is one-half the distance between the two nuclei in two adjacent metal atoms or in a diatomic molecule .



Variation in physical prosperities

Atomic Radius



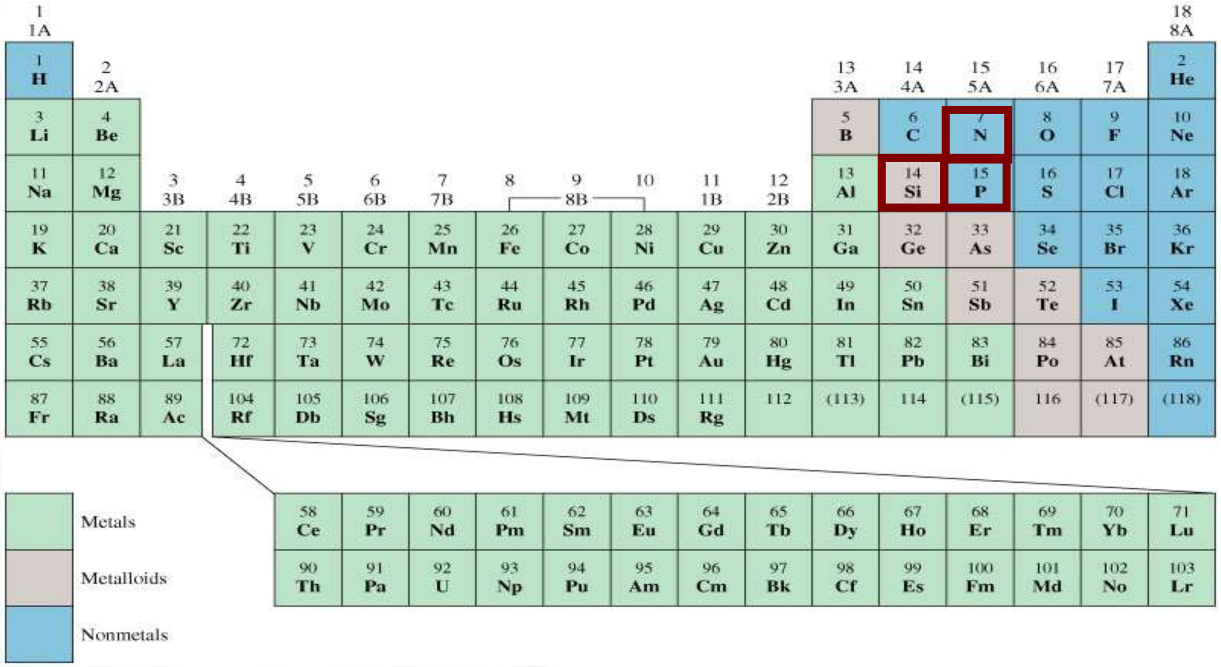
Variation in physical prosperities

Atomic Radius

Example

Referring to a periodic table, arrange the following atoms in order of increasing atomic radius: P, Si, N?

N <P <Si



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1 H																	2 He
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19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
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				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals

▨ Metalloids

🟦 Nonmetals

Variation in physical prosperities

Atomic Radius

Example

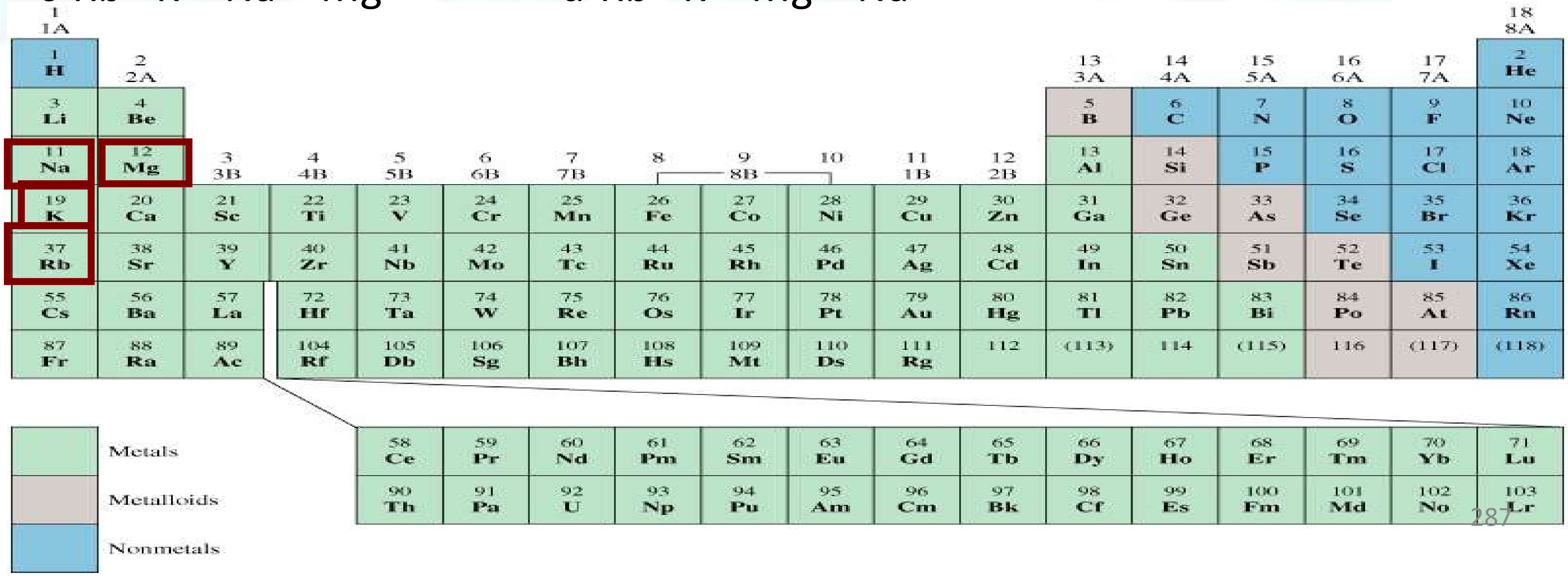
Which choice below correctly lists the elements in order of increasing atomic radius?

a-Na < Mg < K < Rb

b-Mg < Na < K < Rb

c-Rb< K < Na < Mg

d-Rb< K < Mg < Na



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1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals

▨ Metalloids

🟦 Nonmetals

Variation in physical prosperities

Ionic Radius

Ionic Radius: is the radius of a cation or an anion.

If the atom forms an anion, its size (or radius) increases, because the nuclear charge remains the same but the repulsion resulting from the additional electron(s) enlarges the domain of the electron cloud

However, If the atom forms an cation, its size (or radius) decreases, because the nuclear charge remains the same but electron-electron repulsion decreases so the electron cloud shrinks.

– Anion is always larger than atom from which its formed.

– Cation is always smaller than atom from which its formed.

Cations< Anions

Variation in physical prosperities

Ionic Radius

- The ionic radius increases from the top to the bottom of the group
- For ions in different groups: they should be isoelectronic
- Isolelectronic ions:
  - cations < anions: example $Na^+ < F^-$
  - Isolelectronic cations: example Al^{3+}, Mg^{2+}, Na^+

    $Al^{3+} < Mg^{2+} < Na^+$
  - Isoelectronic anoins: example O^{2-}, F^-

    $F^- < O^{2-}$

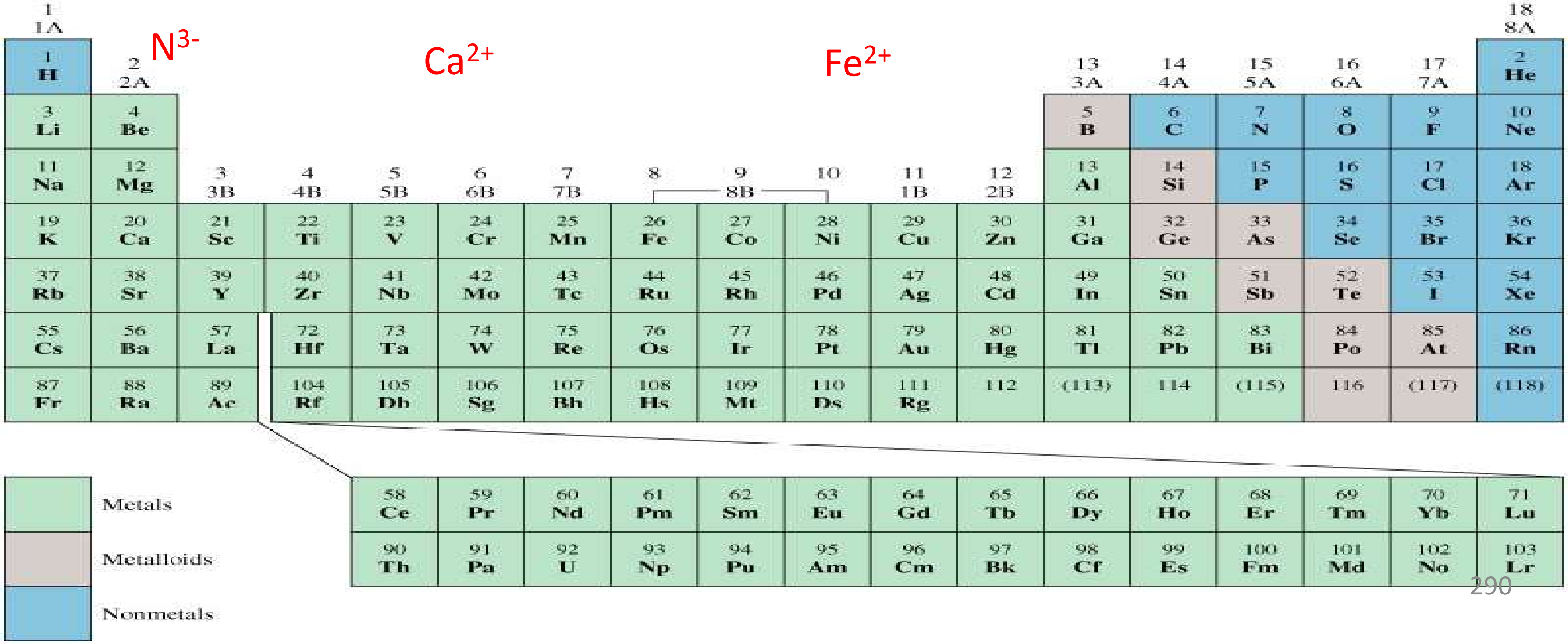
Variation in physical prosperities

Ionic Radius

Example

For each of the following pairs, indicate which one is larger:

a. N^{3-} or F^{-} b- Mg^{2+} or Ca^{2+} c-Fe^{2+} or Fe^{3+}



N^{3-} Ca^{2+} Fe^{2+}

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals
▨ Metalloids
🟦 Nonmetals

Variation in physical prosperities

Ionic Radius

- **Isoelectronic cations:**

➢ **Example:** $^{10}Al^{+3}$, $^{10}Mg^{+2}$, $^{10}Na^{+}$

➢ Arrangement of increasing ionic radius:

$$^{10}Al^{+3} < ^{10}Mg^{+2} < ^{10}Na^{+}$$

- **Isoelectronic anions**

➢ **Example:** $^{10}F^{-}$, $^{10}O^{-2}$, $^{10}N^{-3}$

➢ Arrangement of increasing ionic radius:

$$^{10}F^{-} < ^{10}O^{-2} < ^{10}N^{-3}$$

Variation in physical prosperities

Ionic Radius

Example

Order the following according to the increase in atomic/ionic radius.

N^{3-} Li^{+} C O^{2-}

a- $C < Li^{+} < O^{2-} < N^{3-}$

b- $N^{3-} < O^{2-} < C < Li^{+}$

c- $Li^{+} < C < N^{3-} < O^{2-}$

d- $Li^{+} < C < N^{3-} < O^{2-}$

e- $Li^{+} < C < O^{2-} < N^{3-}$

Always

Cation < neutral < anion

For cation the larger the charge the smaller the radius

For anion the smaller the charge the smaller the radius

Ionization Energy

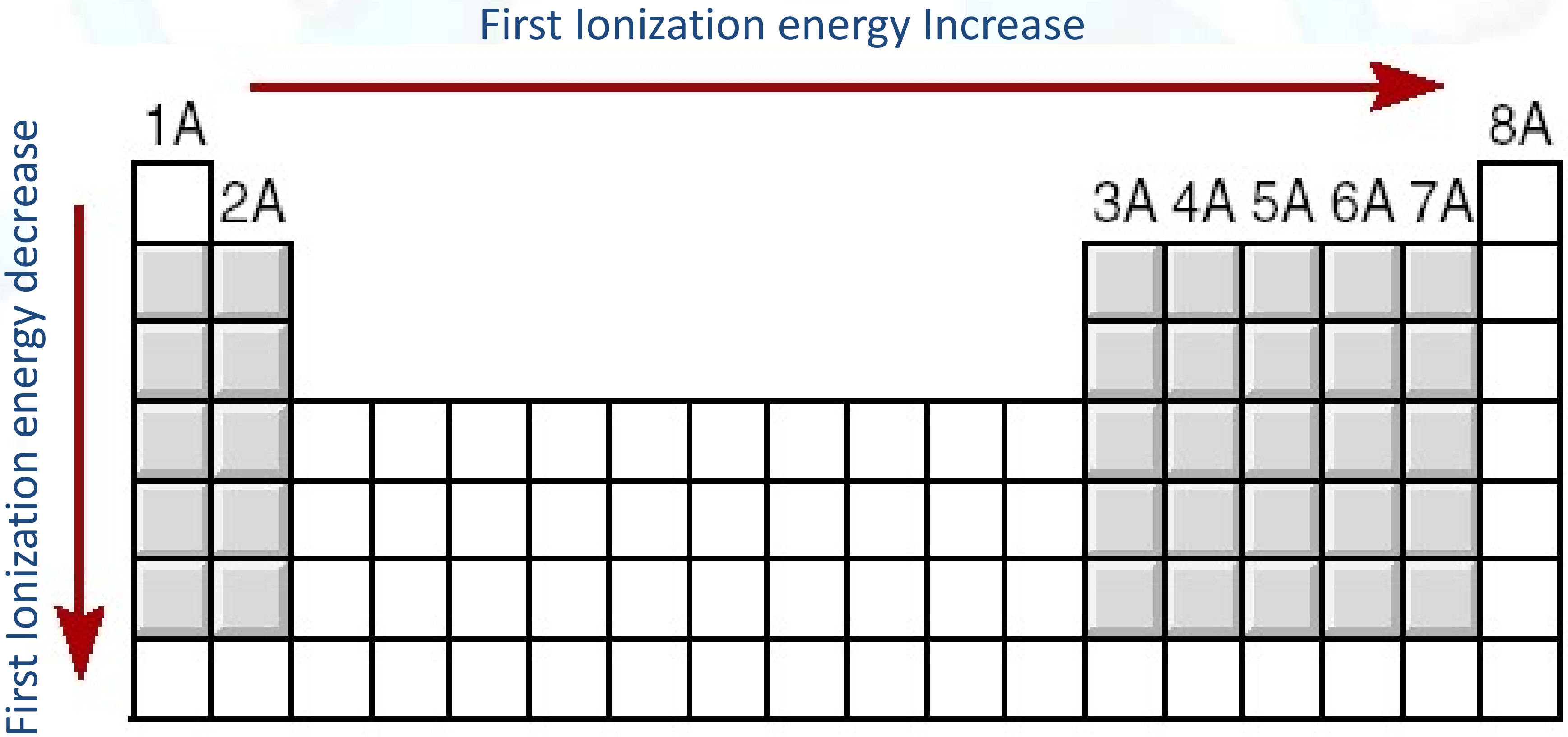
Ionization Energy : is the minimum energy (in kJ/mol) required to remove an electron from a gaseous atom in its ground state.

The higher the ionization energy, the more difficult to remove the electron from the atom.

For a many-electron atom, the amount of energy required to remove the first electron from the atom in its ground state, is called the first ionization energy(I_1). To remove the second electron is called he second ionization energy (I_2) and to remove the third electron is called the third ionization energy (I_3).

$$I_1 < I_2 < I_3 \ldots\ldots$$

Ionization Energy



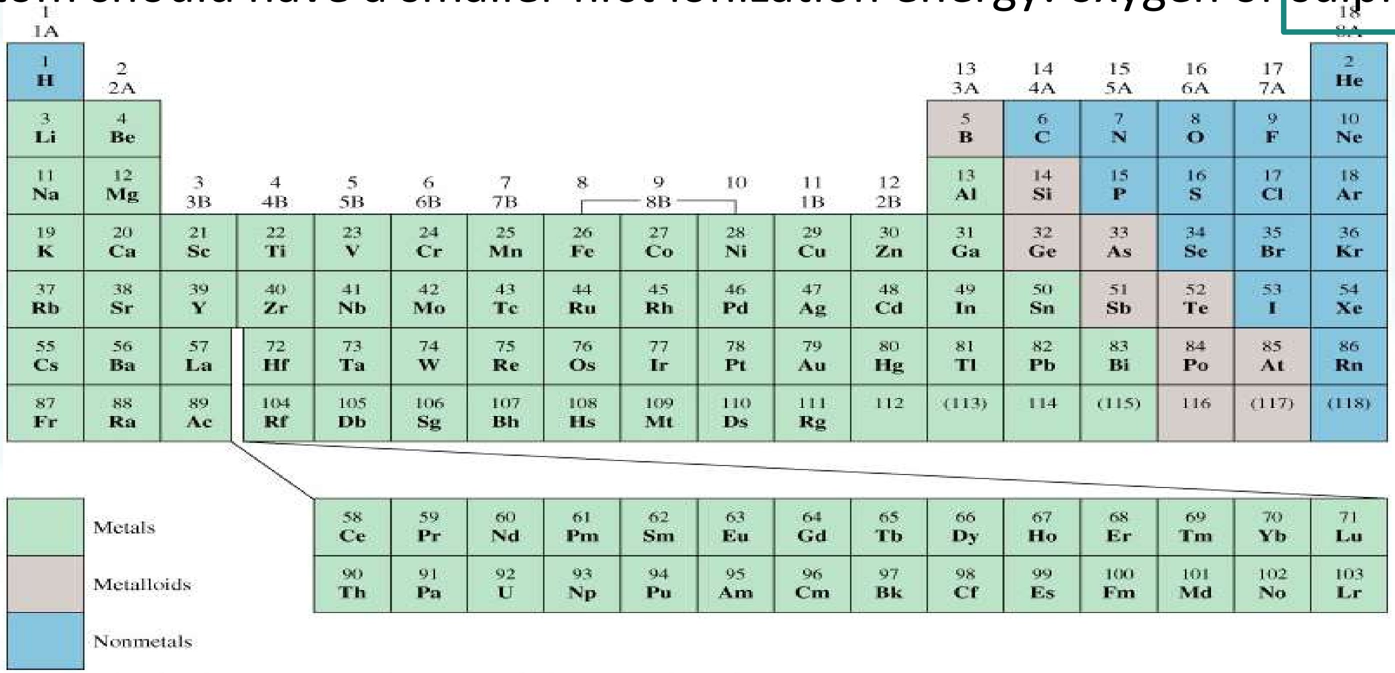
Ionization Energy

Exception

- 2A>3A
- 5A>6A

Example

Which atom should have a smaller first ionization energy: oxygen or sulphur?



1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

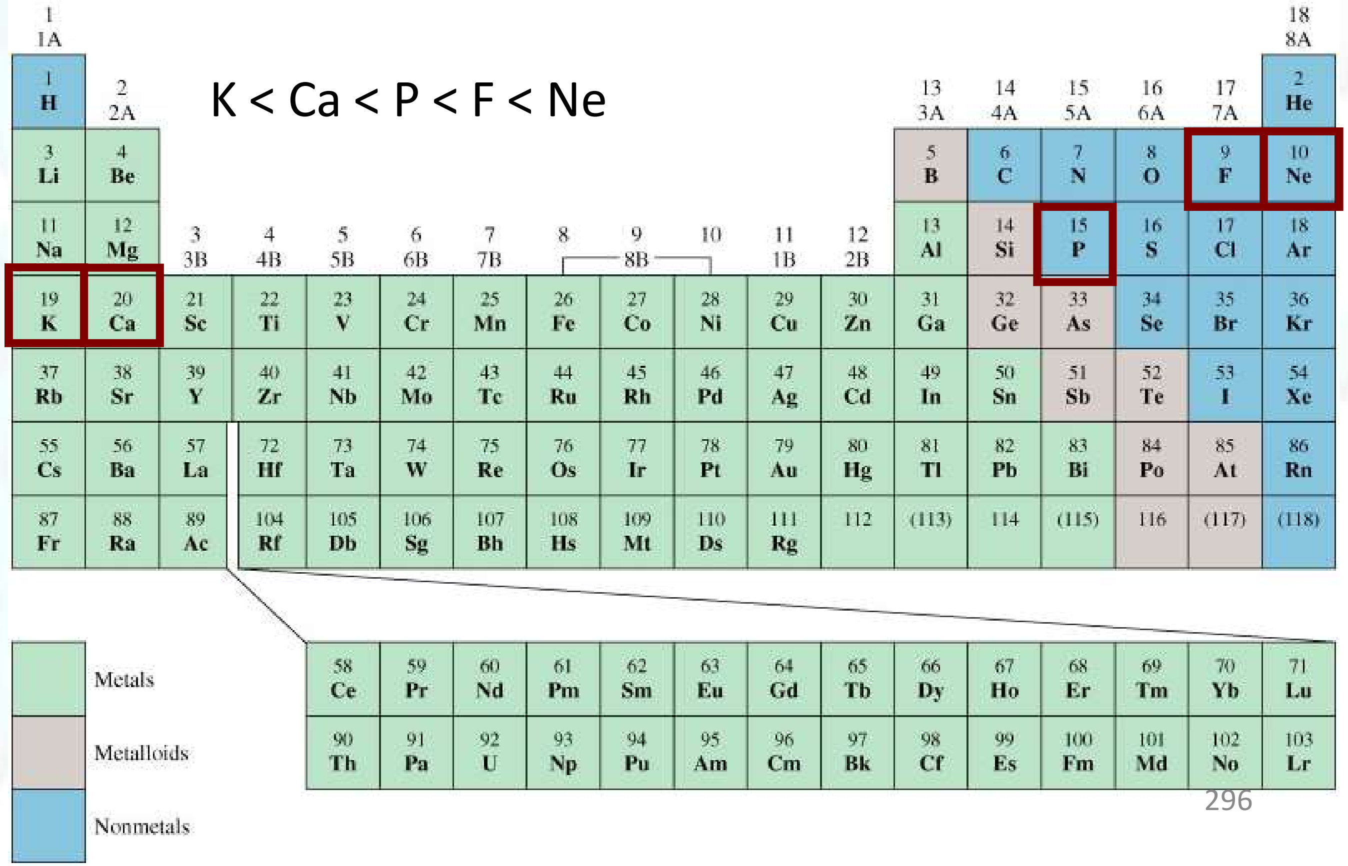
🟩	Metals
⬜	Metalloids
🟦	Nonmetals

Ionization Energy

Example

Arrange the following in order of increasing first ionization energy:

F, K, P, Ca, and Ne.



K < Ca < P < F < Ne

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H**																	2 **He**
3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
55 **Cs**	56 **Ba**	57 **La**	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
87 **Fr**	88 **Ra**	89 **Ac**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112	(113)	114	(115)	116	(117)	(118)
				58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
				90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

Metals

Metalloids

Nonmetals

Electron Affinity

Electron Affinity : is the negative change of the energy that occurs when an electron is accepted by an atom in the gaseous state to form an anion. The higher electron affinity, the greater affinity to accept the electron. The Electron Affinity for non-metal is higher than metal , EA for metalloid fall between metals and nonmetals.

The Electron Affinity decreases from top to the bottom of the group.

The Electron Affinity increase from left to right in period.

Exception

2A<1A

5A<4A

Noble gases have the lowest electron affinities

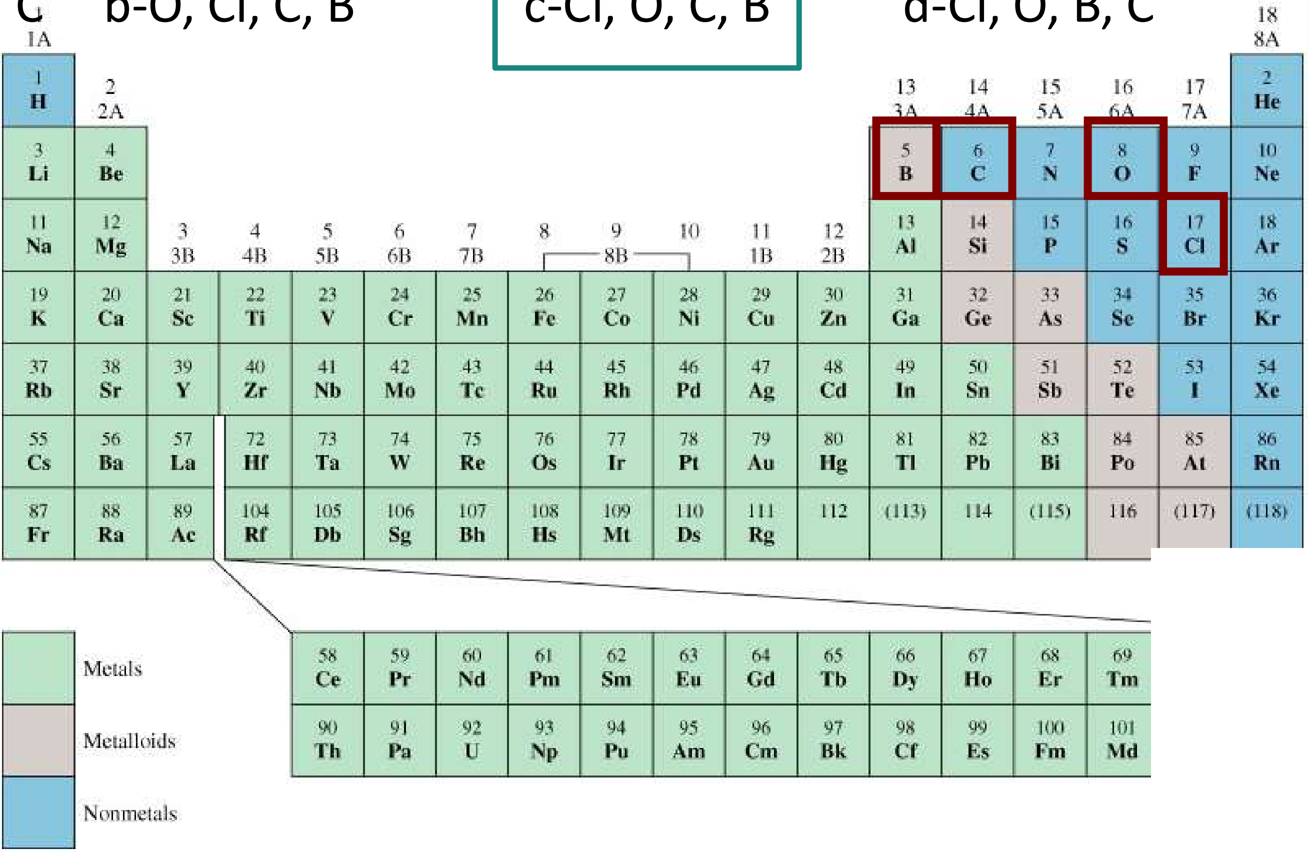
Halogens have the largest electron affinities

Electron Affinity

Example

Which choice correctly lists the elements in order of decreasing electron affinity?

a-O, Cl, B, C b-O, Cl, C, B c-Cl, O, C, B d-Cl, O, B, C



1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm		
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md		

🟩 Metals

⬜ Metalloids

🟦 Nonmetals

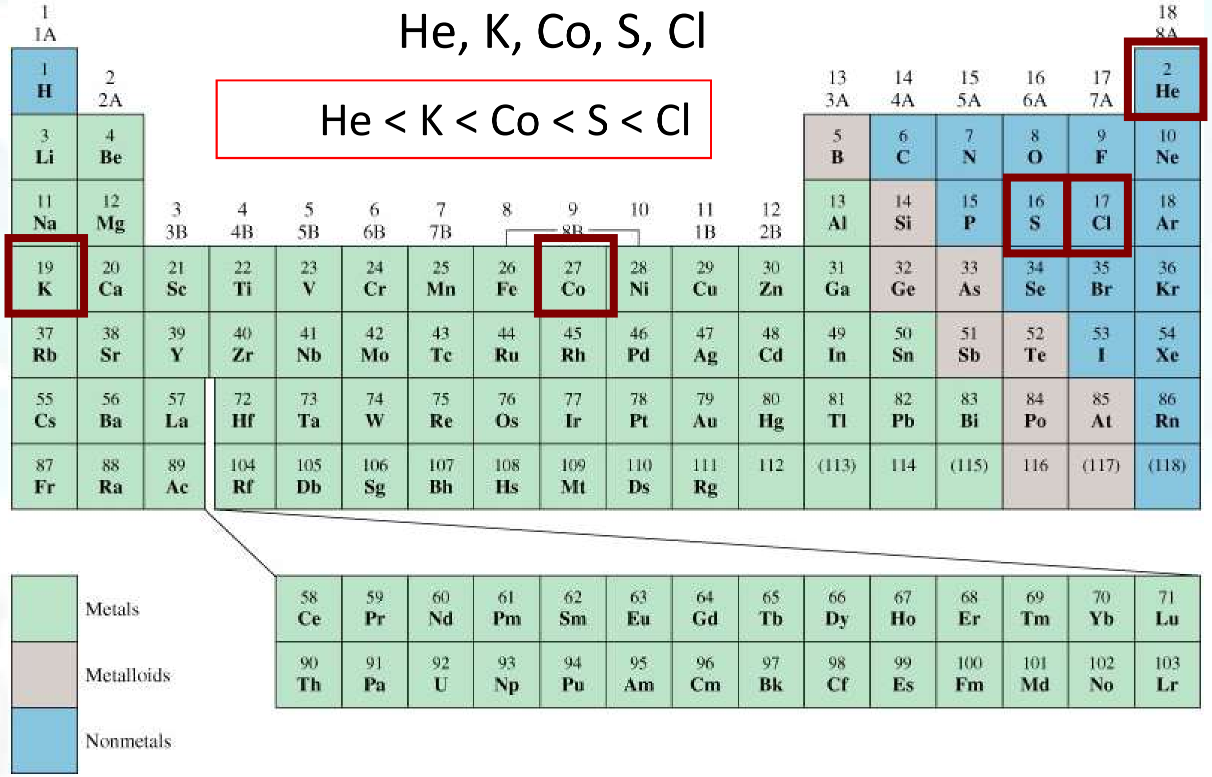
Electron Affinity

Example

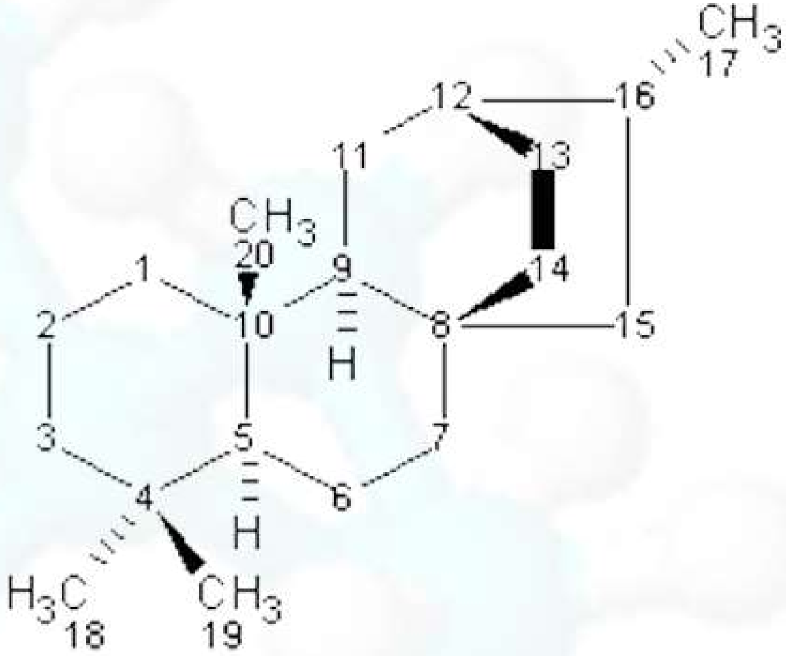
Specify which of the following elements you would expect to have the greatest electron affinity and which have the least:

He, K, Co, S, Cl

He < K < Co < S < Cl



Thank you



Chapter Nine

Chemical Bonding I

Basic Concept

Valence Electrons

Valence electrons: are the outer shell electrons of an atom. The valence electrons are the electrons that participate in chemical bonding.

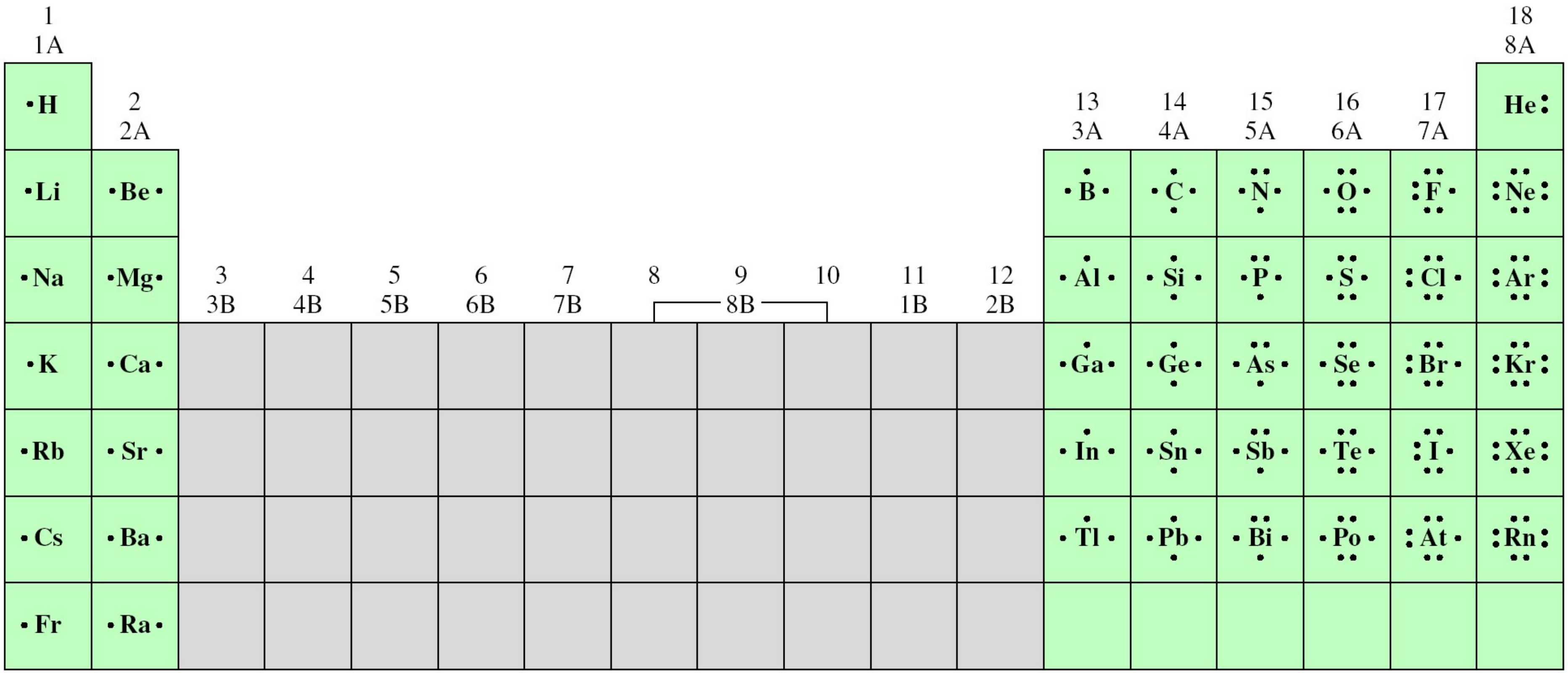
Group	e^- configuration	# of valence e^-
1A	ns^1	1
2A	ns^2	2
3A	ns^2np^1	3
4A	ns^2np^2	4
5A	ns^2np^3	5
6A	ns^2np^4	6
7A	ns^2np^5	7

Lewis Dot Symbols

- Lawis dot symbol consists of the symbol of an element and one dot for each valence electron in an atom of the element.
- The octet rule: in forming chemical bonds, atoms usually gain, lose or share electrons until they have 8 electrons in the outer shell to reach the same electronic configuration of the noble gasses (ns^2 np^6).
- There are two main types of chemical bonds: ionic bond and covalent bond.

Lewis Dot Symbols

Lewis dot symbol for representative elements and noble gases



Ionic Bond

There are two type of bonds: Ionic bond and covalent bond.

An ionic bond is the electrostatic force that holds ions together in an ionic compound.

Ionic bond occurs normally between metal and nonmetal and the electron transfer from element to another element.

Example:

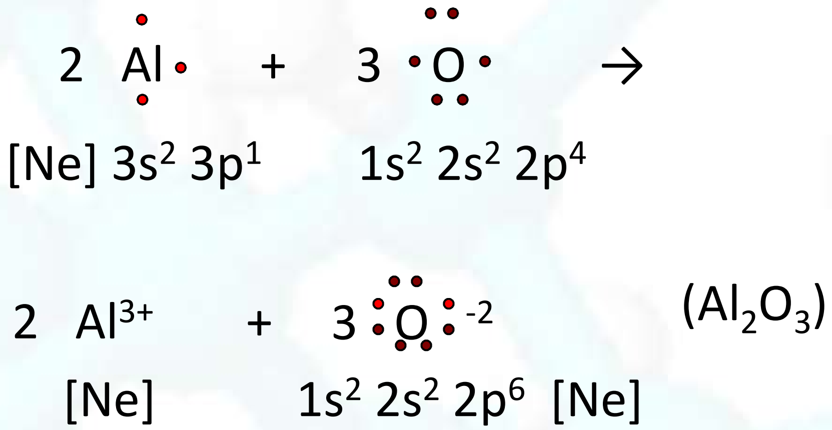
$$Li + F \longrightarrow Li^{+} + F^{-}$$

$1S^2\ 2S^1$ $1S^2 2S^2 2P^5$ $1S^2$ $1S^2 2S^2 2P^6$

Ionic Bond

Example 2

Use Lewis dot symbol to show formation of Al_2O_3



$2\ Al + 3\ O \rightarrow$

$[Ne]\ 3s^2\ 3p^1$ $\quad$ $1s^2\ 2s^2\ 2p^4$

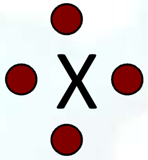
$2\ Al^{3+} + 3\ O^{-2}$ $\quad$ (Al_2O_3)

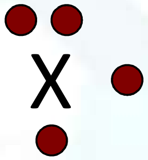
$[Ne]$ $\quad$ $1s^2\ 2s^2\ 2p^6\ [Ne]$

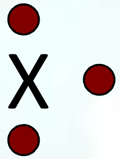
Ionic Bond

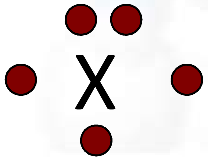
Example:

An atom in the ground state has atomic number Z=5. Choose the correct electron-dot structure which represents this atom?

a.


b.


[✓ c.]


d.


The electron configuration Z =5

$1s^2\ 2s^2\ 2p^1$

Valence electrons are $2s^2\ 2p^1$

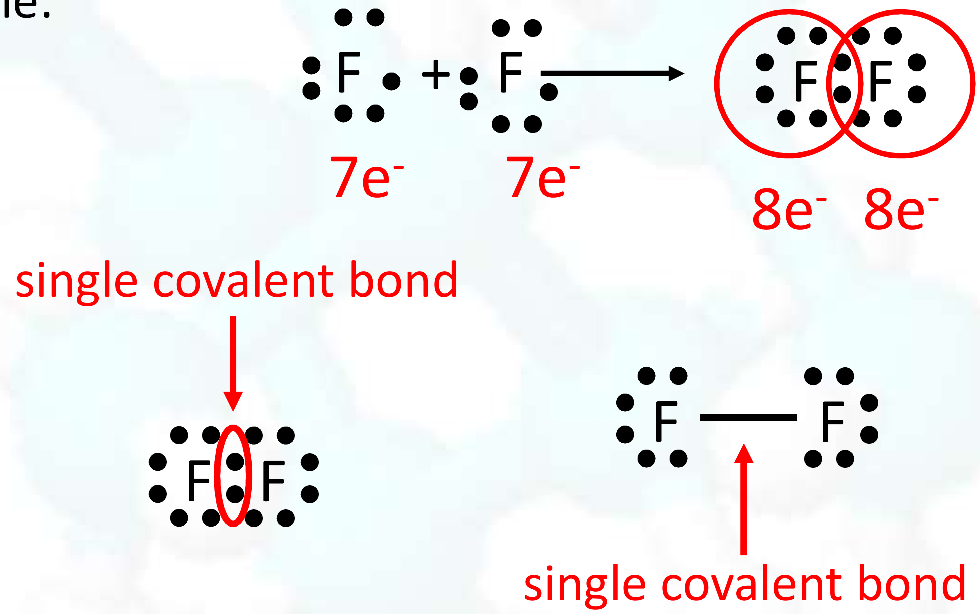
VE = 3 e

Covalent Bond

Covalent bond, is a bond in which two electrons are shared by two atoms.

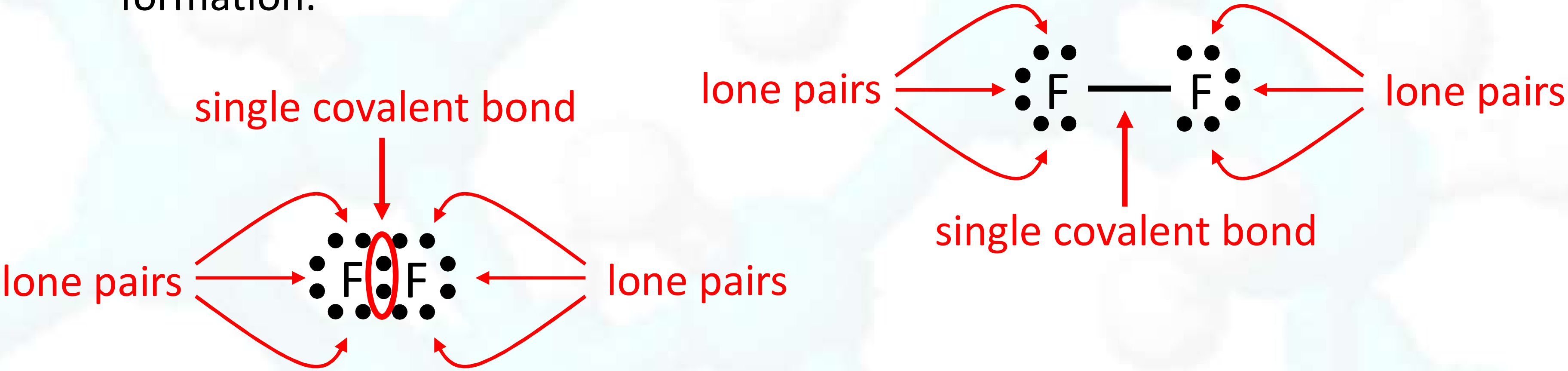
Covalent bond occurs normally between nonmetal and nonmetal and the electron only shared between elements (not transfer from element to another element).

Example:



Covalent Bond

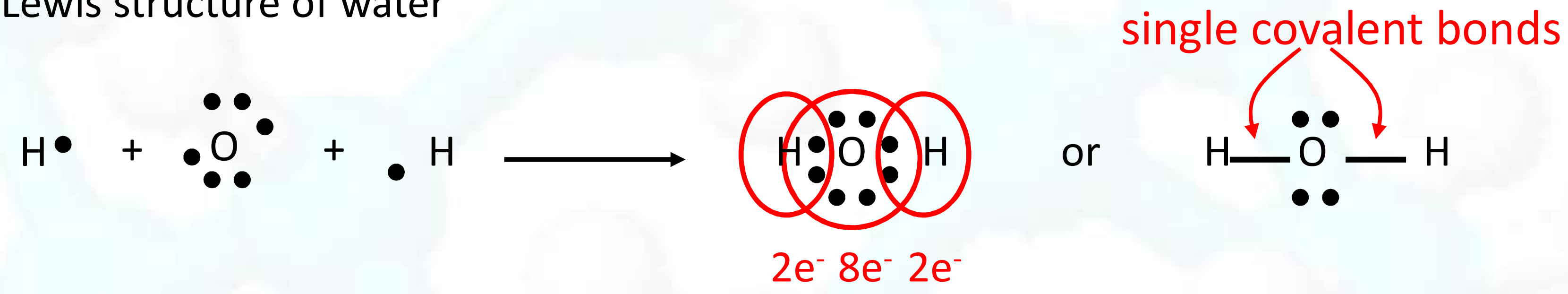
lone pairs: pairs of valence electrons that are not involved in covalent bond formation.



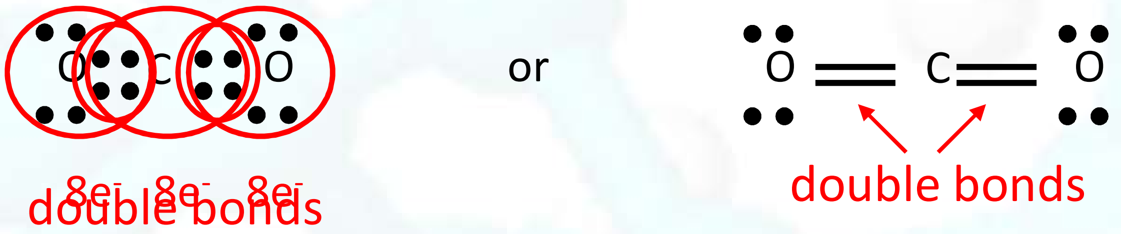
A Lewis structure is a representation of covalent bonding in which shared electron pairs are shown either as lines or as pairs of dots between two atoms, and lone pairs are shown as pairs of dots on individual atoms.

Covalent Bond

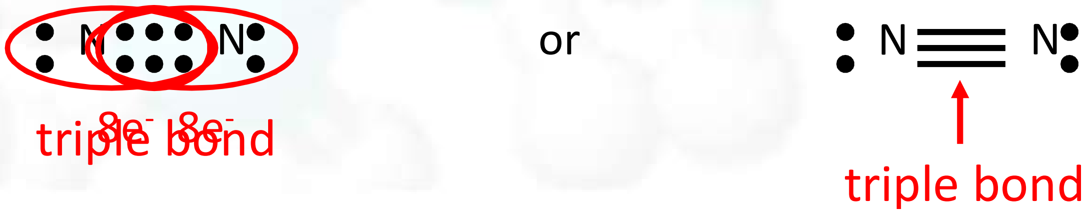
Lewis structure of water



Double bond – two atoms share two pairs of electrons

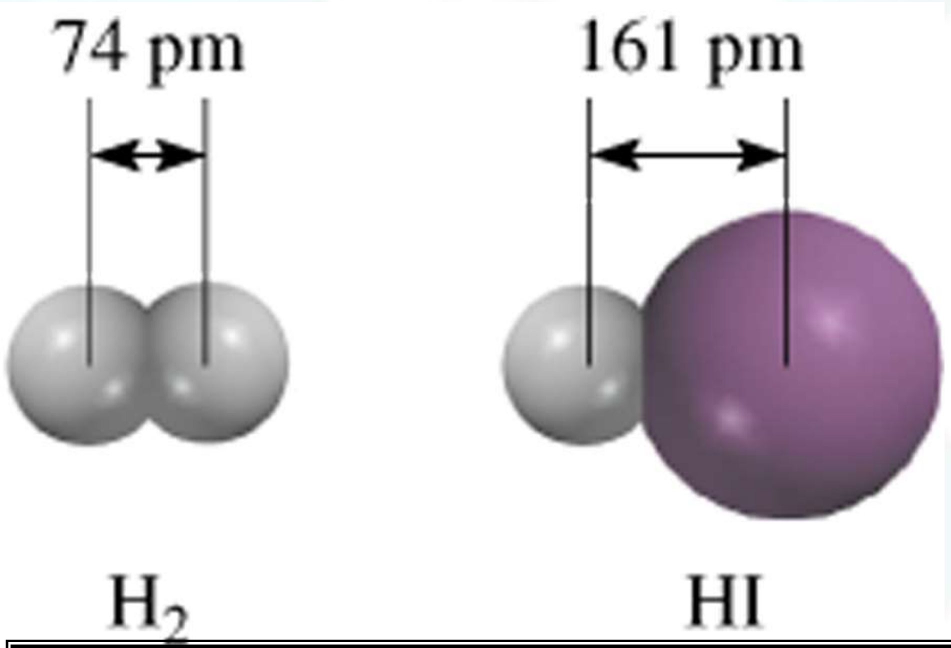


Triple bond – two atoms share three pairs of electrons



Valence Electrons

Bond length is defined as the distance between the nuclei of two covalently bonded atoms in a molecule



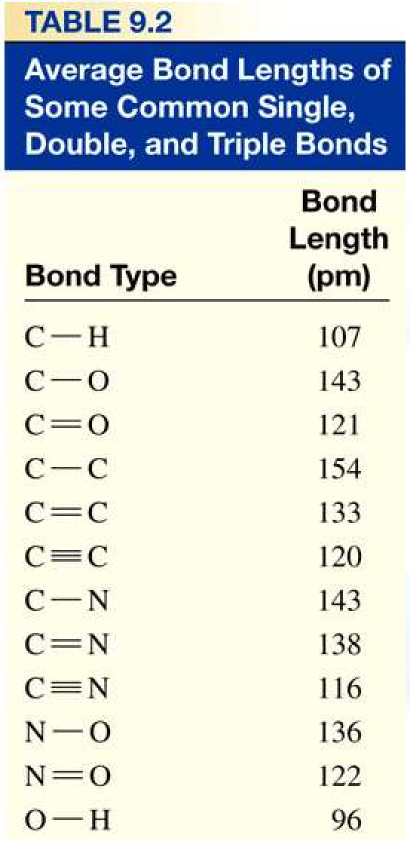


TABLE 9.2

Average Bond Lengths of Some Common Single, Double, and Triple Bonds

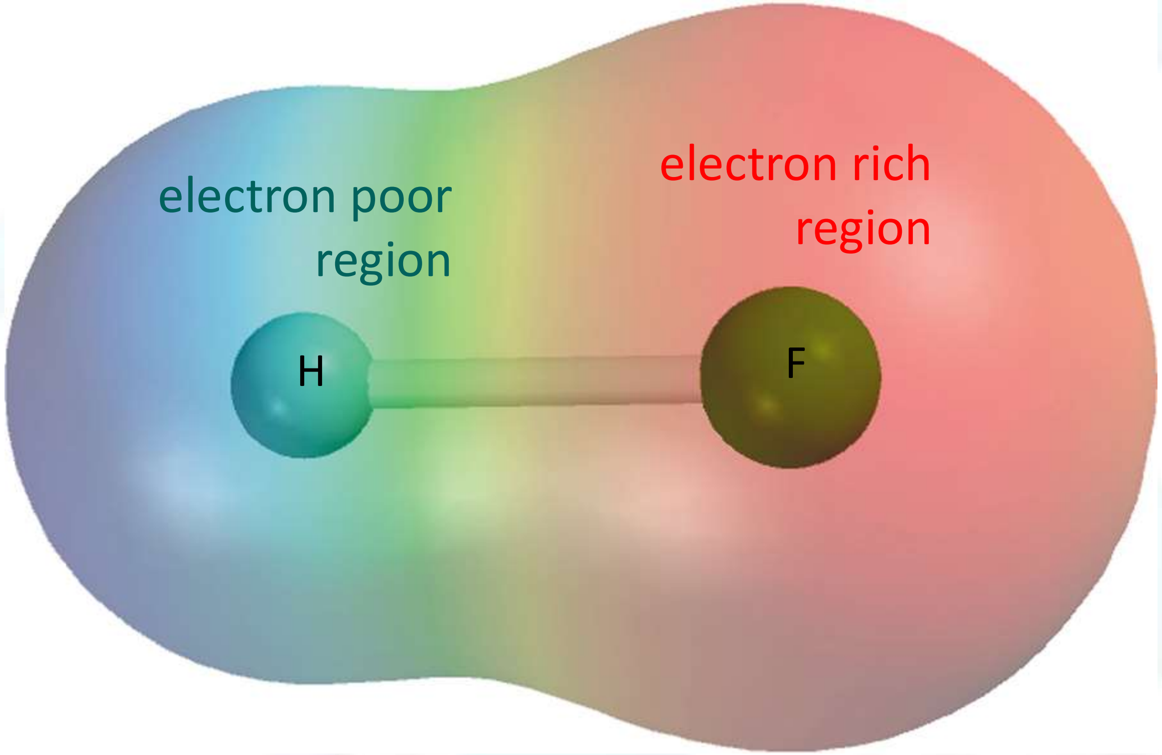
Bond Type	Bond Length (pm)
C—H	107
C—O	143
C=O	121
C—C	154
C=C	133
C≡C	120
C—N	143
C=N	138
C≡N	116
N—O	136
N=O	122
O—H	96

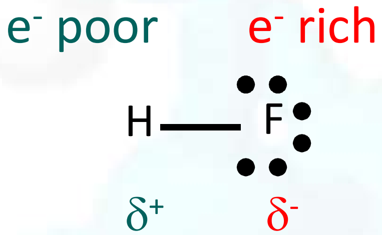
Bond Lengths

Triple bond < Double Bond < Single Bond

Electronegativity

Polar covalent bond or polar bond is a covalent bond with greater electron density around one of the two atoms





Electronegativity

Electronegativity : the ability of an atom to attract toward itself the electrons in a chemical bond.

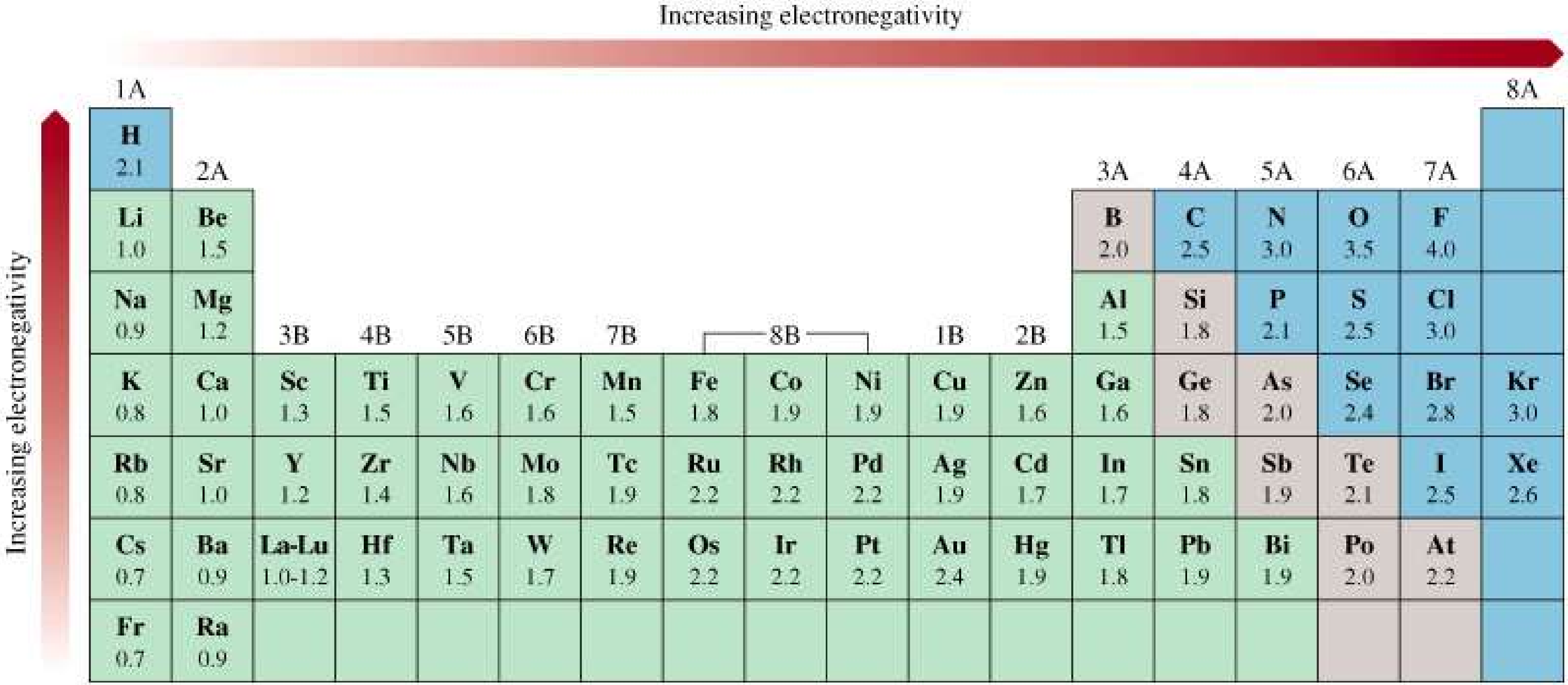
Elements with high electronegativity have a greater tendency to attract electrons than do elements with low electronegativity.

Electron Affinity - measurable, Cl is highest

$$X_{(g)} + e^{-} \longrightarrow X^{-}_{(g)}$$

Electronegativity - relative, F is highest

Electronegativity



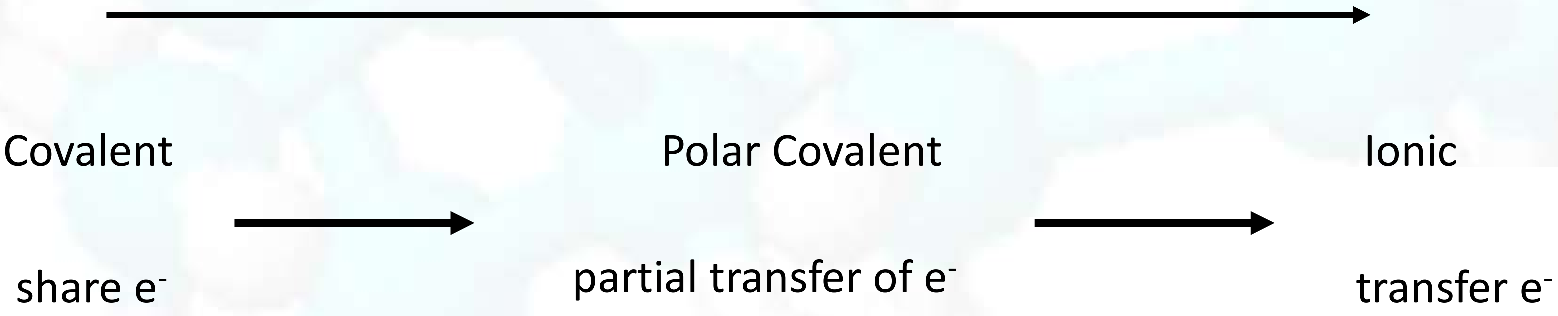
1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H 2.1																	
Li 1.0	**Be** 1.5											**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0	
Na 0.9	**Mg** 1.2											**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0	
K 0.8	**Ca** 1.0	**Sc** 1.3	**Ti** 1.5	**V** 1.6	**Cr** 1.6	**Mn** 1.5	**Fe** 1.8	**Co** 1.9	**Ni** 1.9	**Cu** 1.9	**Zn** 1.6	**Ga** 1.6	**Ge** 1.8	**As** 2.0	**Se** 2.4	**Br** 2.8	**Kr** 3.0
Rb 0.8	**Sr** 1.0	**Y** 1.2	**Zr** 1.4	**Nb** 1.6	**Mo** 1.8	**Tc** 1.9	**Ru** 2.2	**Rh** 2.2	**Pd** 2.2	**Ag** 1.9	**Cd** 1.7	**In** 1.7	**Sn** 1.8	**Sb** 1.9	**Te** 2.1	**I** 2.5	**Xe** 2.6
Cs 0.7	**Ba** 0.9	**La-Lu** 1.0-1.2	**Hf** 1.3	**Ta** 1.5	**W** 1.7	**Re** 1.9	**Os** 2.2	**Ir** 2.2	**Pt** 2.2	**Au** 2.4	**Hg** 1.9	**Tl** 1.8	**Pb** 1.9	**Bi** 1.9	**Po** 2.0	**At** 2.2	
Fr 0.7	**Ra** 0.9																

Electronegativity

Classification of bonds by difference in electronegativity

Difference	Bond Type
0	Covalent
≥ 2	Ionic
0 < and <2	Polar Covalent

Increasing difference in electronegativity



Covalent → Polar Covalent → Ionic

share e^- → partial transfer of e^- → transfer e^-

Electronegativity

Example

Classify the following bonds as ionic, polar covalent, or covalent

A) HCl 3-2.1 =0.9

Polar covalent

b) KF 4-0.8 =3.2

Ionic

c) C-C 2.5-2.5 =0

covalent

Electronegativity

Example

Classify the following bonds as ionic, polar covalent, or covalent

A) CsCl 3 – 1 = 2

Ionic

b) H_2S 2.5 - 2.1 = 0.4

Polar covalent

c) N-N 3 - 3 = 0

covalent

Writing Lewis Structures

1. Draw skeletal structure of compound showing which atoms are bonded to each other. Put least electronegative element in the center.
2. Count total number of valence e^-. Add 1 for each negative charge. Subtract 1 for each positive charge.
3. Draw single covalent bond between the central atom and each of the surrounding atom, and complete an octet for all the surrounded atoms ***except*** hydrogen. Count and Compare it with the number of electrons in step 2 if they are identical stop, if it is less add the remaining electrons to the central atom.
4. If still the central atom has no octet, use lone pair/s on the one of the surrounded atom to form double or triple bond with the central atom.

Writing Lewis Structures

Example 1

Write the Lewis structure of nitrogen trifluoride (NF_3).

Step 1 – N is less electronegative than F, put N in center

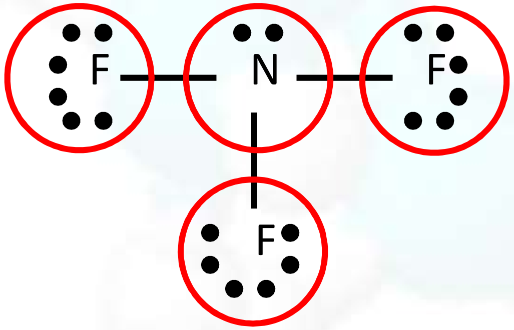
Step 2 – Count valence electrons N - 5 ($2s^22p^3$) and F - 7 ($2s^22p^5$)

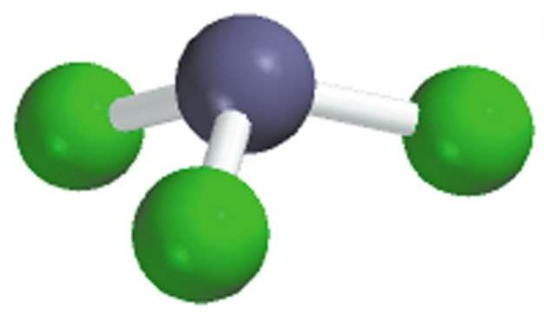
5 + (3 x 7) = 26 valence electrons

Step 3 – Draw single bonds between N and F atoms and complete octets on N and F atoms.

Step 4 - Check, are # of e^- in structure equal to number of valence e^- ?

3 single bonds (3x2) + 10 lone pairs (10x2) = 26 valence electrons

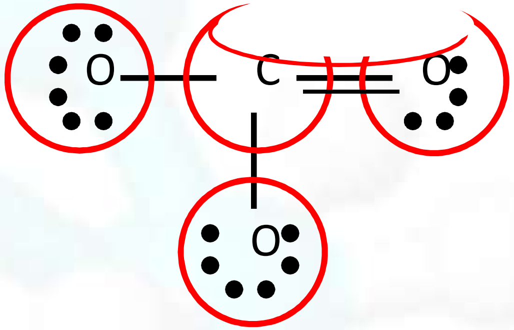




Writing Lewis Structures

Example 2

Write the Lewis structure of the carbonate ion (CO_3^{2-}).



Step 1 – C is less electronegative than O, put C in center

Step 2 – Count valence electrons C - 4 ($2s^22p^2$) and O - 6 ($2s^22p^4$) -2 charge – $2e^-$

4 + (3 x 6) + 2 = 24 valence electrons

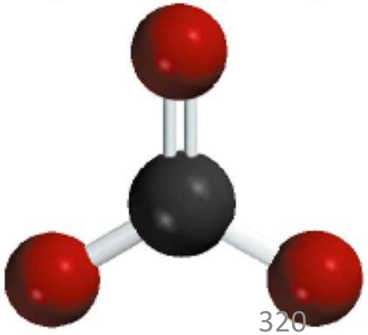
2 single bonds (2x2) = 4
1 double bond = 4
8 lone pairs (8x2) = 16
Total = 24

Step 3 – Draw single bonds between C and O atoms and complete octet on C and O atoms.

Step 4 - Check, are # of e^- in structure equal to number of valence e^- ?

3 single bonds (3x2) + 10 lone pairs (10x2) = 26 valence electrons

Step 5 - Too many electrons, form double bond and re-check # of e^-



Writing Lewis Structures

Example 3

Write the Lewis structure for nitric acid (HNO_3) in which the three O atoms are bonded to the central N and H atom is bonded to one of the O atoms?

Step 1: put N in center ,surrounded by 3O atoms , H bonded to one of the O.

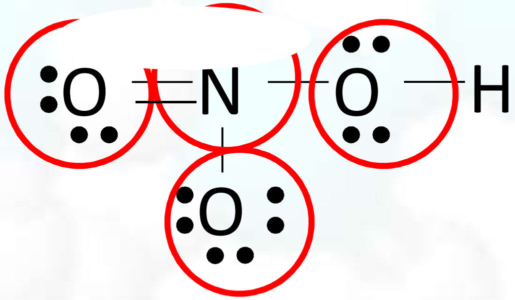
Step 2: Count the valence electrons 5 + (3 x 6) +1 = 24 valence e^-

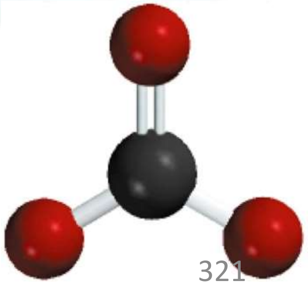
Step 3: Draw single bonds between N and O atoms and O and H complete octet on O and N atoms. Use all the valence e (step 2)

Step 4: Check, are # of e^- in structure equal to number of valence e^- ?

4 single bonds (4x2) + 9 lone pairs (9x2) = 26 valence electrons

Step 5 - Too many electrons, form double bond and re-check # of e^-





Formal Charges and Lewis Structure

formal charge is the difference between the number of valence electrons in an isolated atom and the number of electrons assigned to that atom in a Lewis structure.

formal charge on an atom in a Lewis structure = total number of valence electrons in the free atom - total number of nonbonding electrons - total number of bond

The sum of the formal charges of the atoms in a molecule or ion must equal the charge on the molecule or ion.

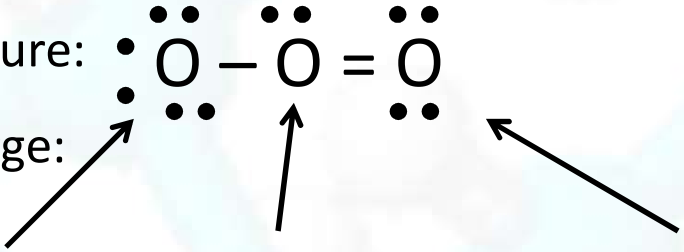
Formal Charges and Lewis Structure

Example:

Ozone molecule (O_3)

Lewis Structure:

Formal Charge:



6 - 6 - 1 = -1

6 - 2 - 3 = +1

6 - 4 - 2 = 0

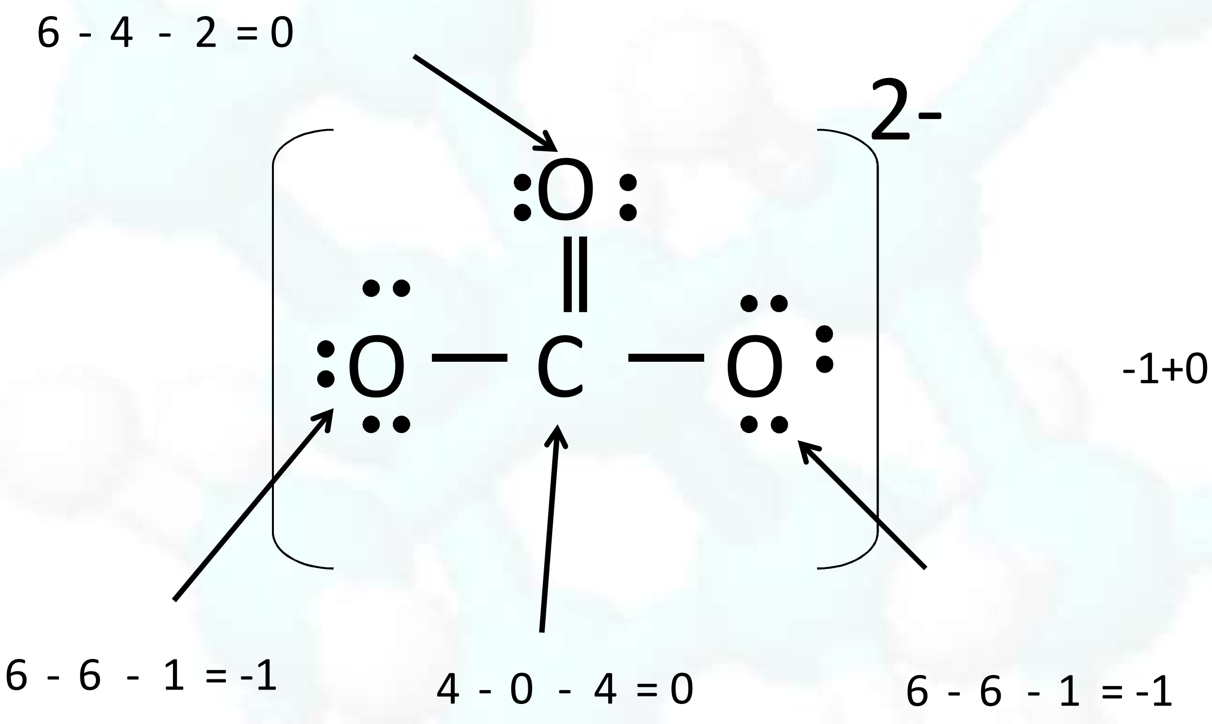
-1+1+0=0

formal charge on an atom in a Lewis structure = total number of valence electrons in the free atom - total number of nonbonding electrons - total number of bond

Formal Charges and Lewis Structure

Example 2

Write the formal charges for the carbonate ion (CO_3^{2-})



-1+0 -1+0=-2

Formal Charges and Lewis Structure

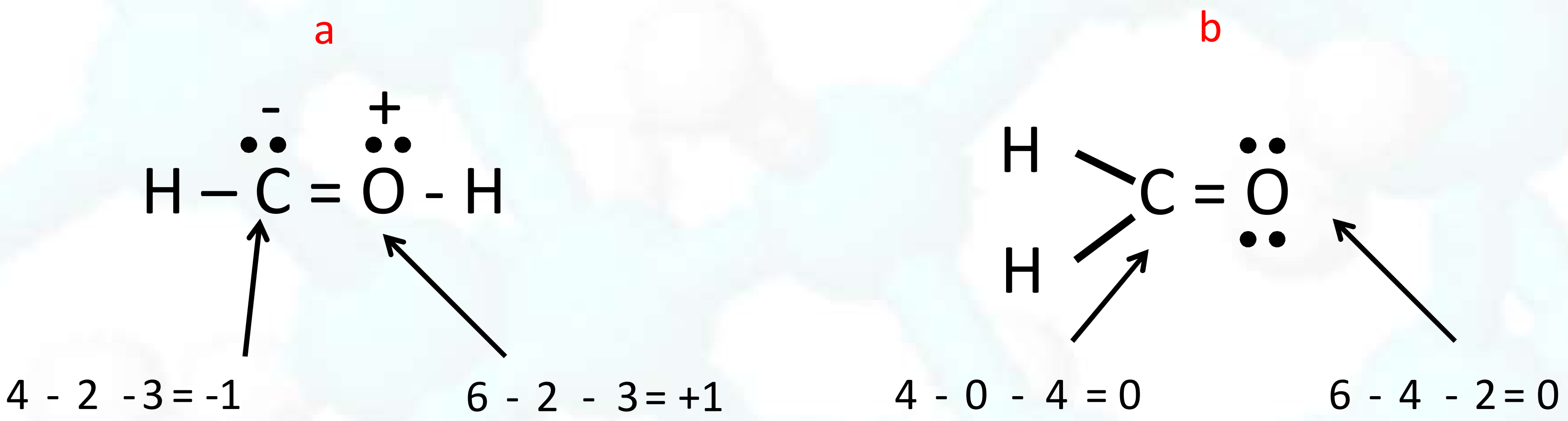
Sometimes there is more than one acceptable Lewis structure for a given species. In such cases, we can often select the most plausible Lewis structure by using formal charges and the following guidelines:

1. For molecules, a Lewis structure in which there are no formal charges is preferable to one in which formal charges are present.
2. Lewis structures with large formal charges (+2, +3, and/or -2, -3, and so on) are less plausible than those with small formal charges.
3. Among Lewis structures having similar distributions of formal charges, the most plausible structure is the one in which negative formal charges are placed on the more electronegative atoms.

Formal Charges and Lewis Structure

Example

Draw the most likely Lewis structure for formaldehyde (CH_2O).



4 - 2 -3 = -1

6 - 2 - 3 = +1

4 - 0 - 4 = 0

6 - 4 - 2 = 0

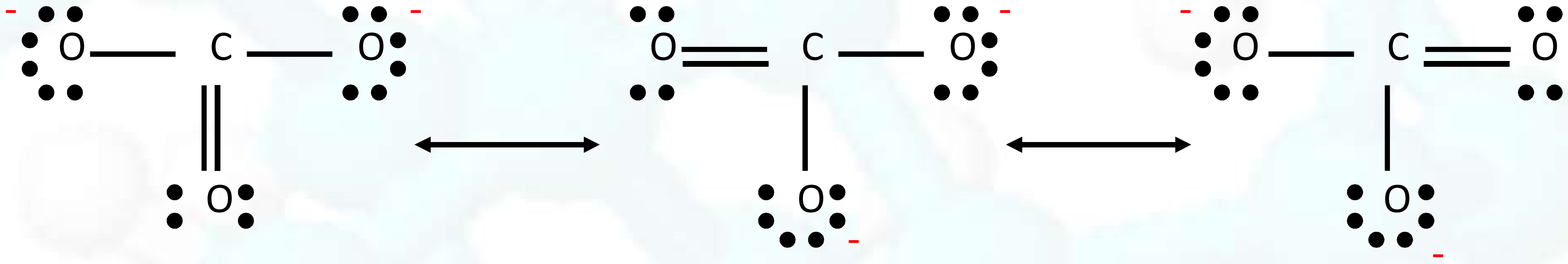
(b) is the more likely structure
because it carries no formal charges

The Concept of Resonance

A resonance structure is one of two or more Lewis structures for a single molecule that cannot be represented accurately by only one Lewis structure.

Example

What are the resonance structures of the carbonate $(CO_3)^{-2}$ ion?



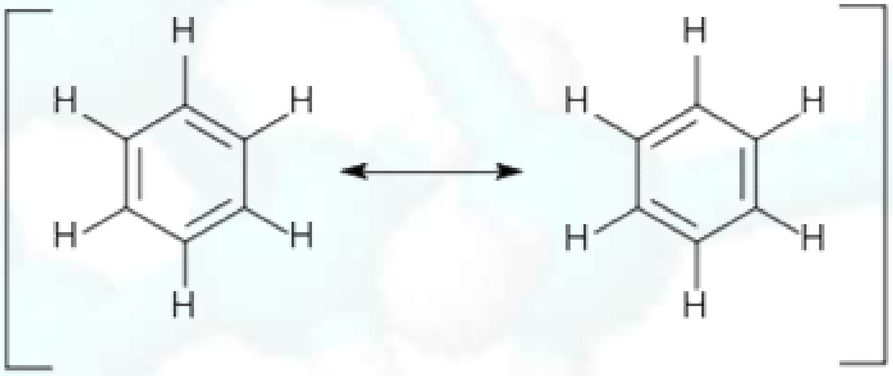
The Concept of Resonance

Example

What are the resonance structures of Ozon (O_3)?

Example

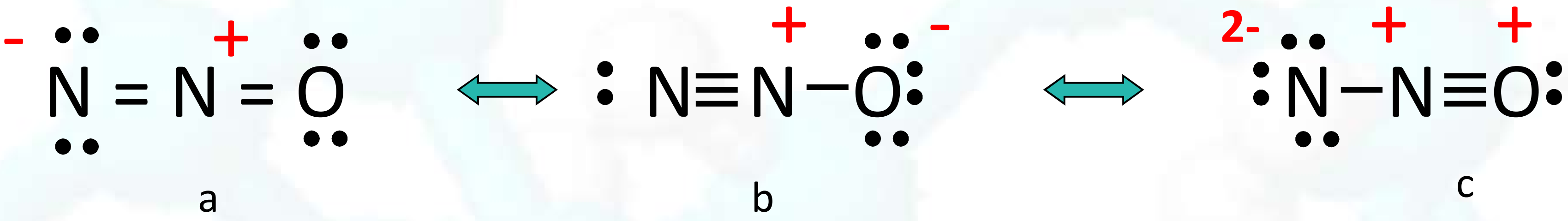
What are the resonance structures of benzene molecule (C_6H_6)?



The Concept of Resonance

Example

Draw three resonance structures for the molecule nitrous oxide, N_2O (the atomic arrangement is NNO). Indicate formal charges. Rank the structures



Structure (b) is the most important one because the negative charge is on the more electronegative oxygen atom. Structure (c) is the least important one because it has a larger separation of formal charges. Also, the positive charge is on the more electronegative oxygen atom.

Exceptions to the Octet Rule

Incomplete octet

The number of electrons surrounding the central atom in stable molecule is less than 8.

Example

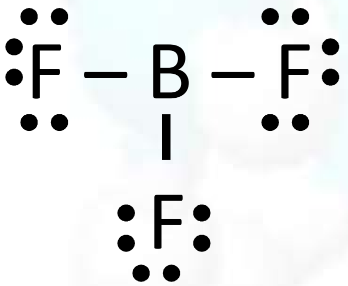
BeH_2

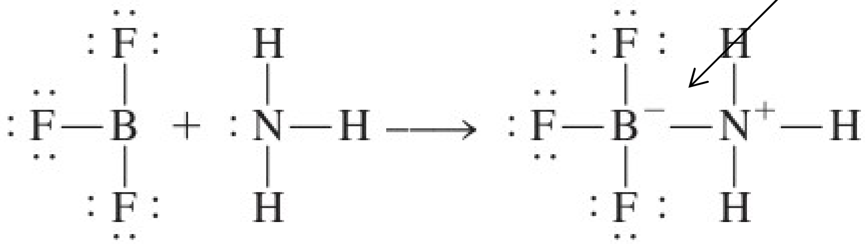
H —— Be —— H

coordinate covalent bond (also referred to as a dative bond), defined as a covalent bond in which one of the atoms donates both electrons.

The electron configuration of beryllium is 1 s^2 $2s^2$; it has two valence electrons in the 2s orbital.

Also group 3A (Al, B).





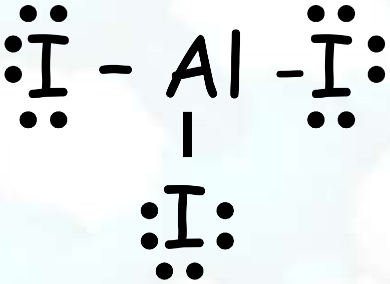
Exceptions to the Octet Rule

Incomplete octet

Example

Draw the Lewis structure for aluminium triiodide (AlI_3)

The outer-shell electron configurations of Al and I are $3s^2$ $2p^1$ and $5s^2$ $5p^5$ respectively. The total number of valence electrons is 3 + 3 x 7 or 24. Because Al is less electronegative than I, it occupies a central position and forms three bonds with the I atoms:



there are only six valence electrons around the Al atom. Thus, AlI_3 is an example of the incomplete octet.

Exceptions to the Octet Rule

Odd-Electron Molecules

molecules contain an odd number of electrons also called radical.

Radical: atom has one electron alone

Example are nitric oxide (NO) and nitrogen dioxide (NO_2)

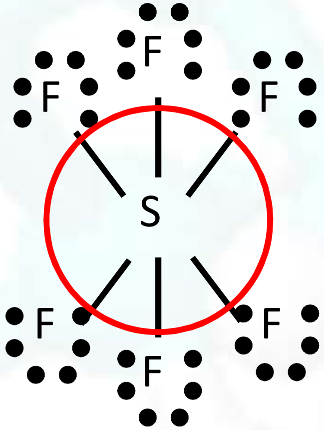
Exceptions to the Octet Rule

The Expanded Octet

3rd period and beyond (4th, 5th, 6th, 7th) may form molecules in which the central atom is surrounded by more than 8 electrons

Example

The electron configuration of sulfur is [Ne]$3s^2$ $3p^4$. In SF_6 , each of sulfur's six valence electrons forms a covalent bond with a fluorine atom, so there are 12 electrons around the central sulfur atom:



Exceptions to the Octet Rule

The Expanded Octet

Example

Draw Lewis structure for phosphorus pentaflouridde (PF_5), in which all five F atoms are bounded to the central P atom.?

Valance electrons = 5 + (7 x 5) = 40 e

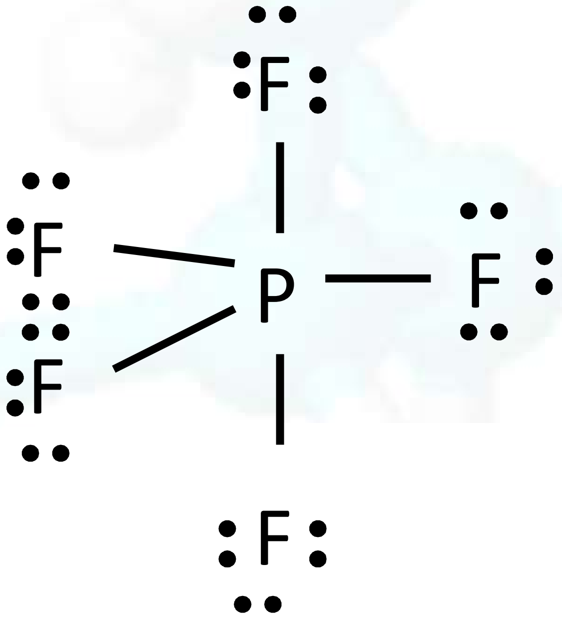
5x6=30 electron (lone pair)

5x2=10 electron (bond)

Total is 40 electron

P has 10 electron (5x2)

Thus it expanded octet

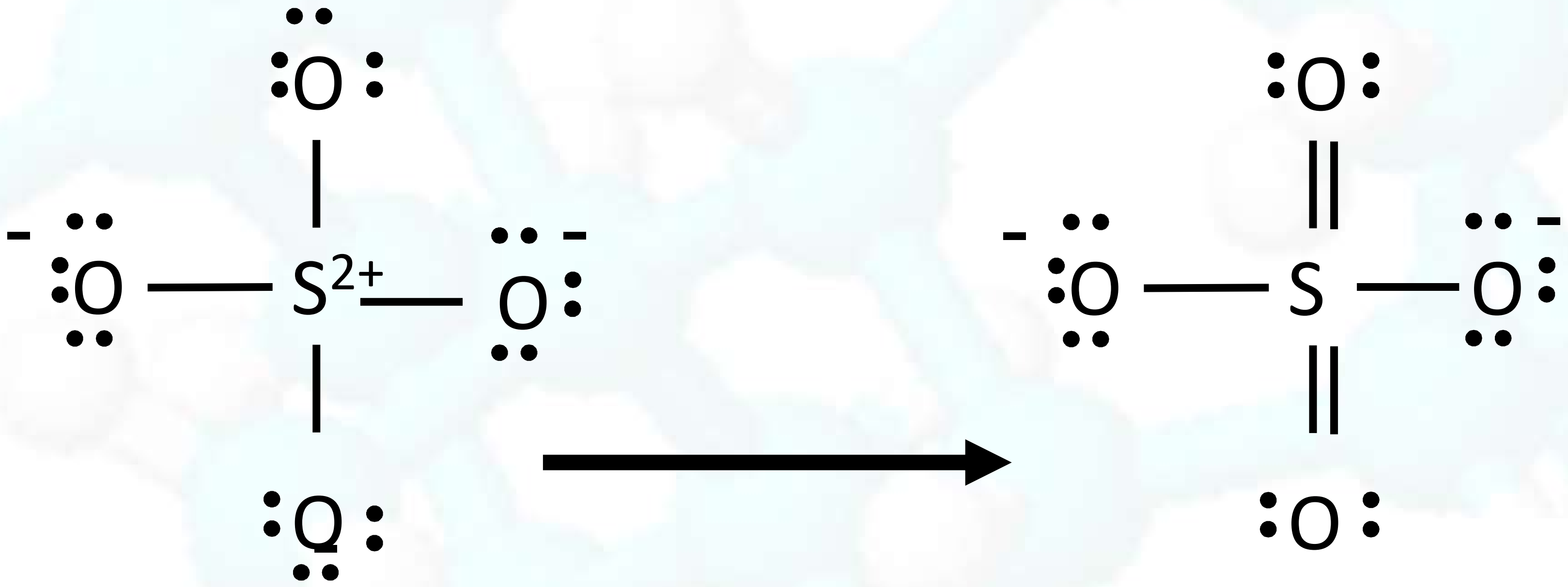


Exceptions to the Octet Rule

Exceptions to the Octet Rule

Example

Draw Lewis structure for sulphate ion (SO_4^{2-}), in which all four O atoms are bounded to the central S atom.



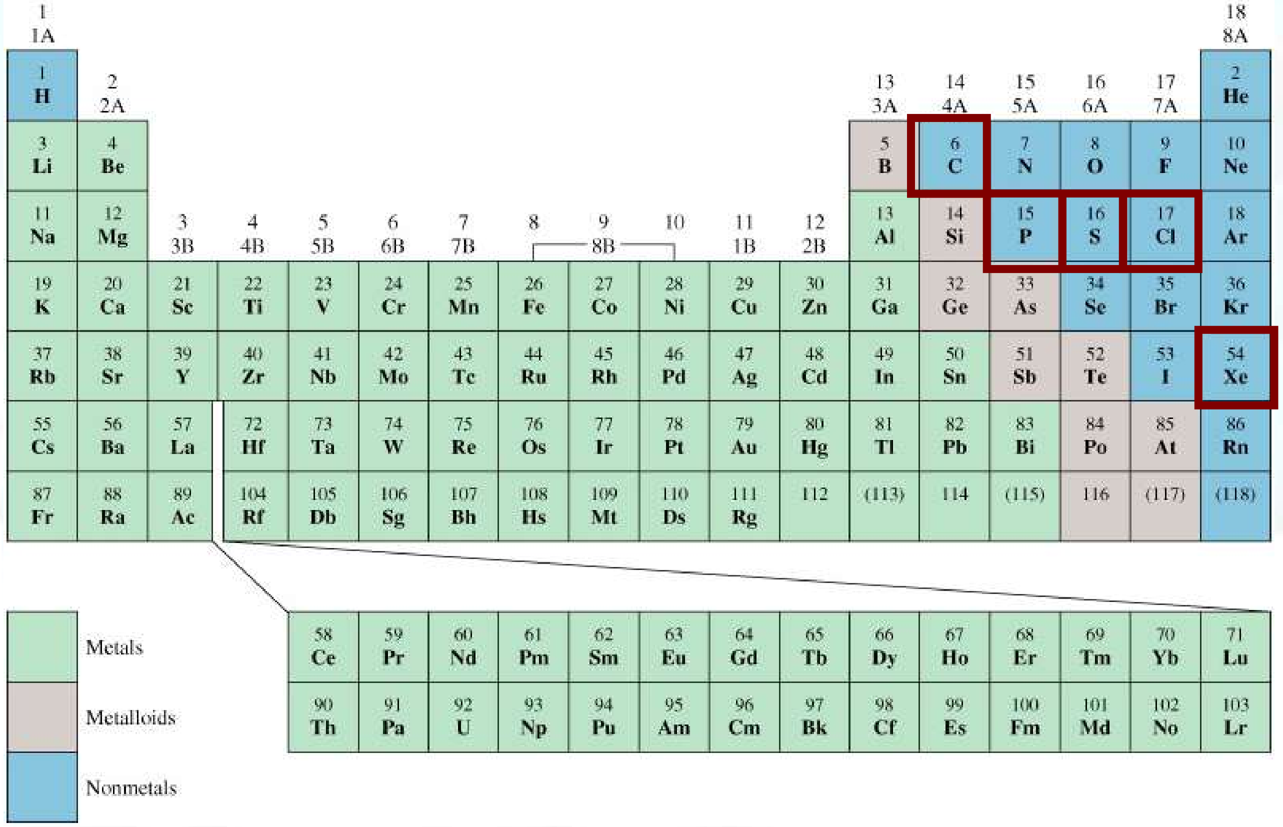
This structure has an expanded octet on S, BUT it is more correct because, it has less formal charges.

Exceptions to the Octet Rule

Example

When acting as a central atom, which of the following cannot expand its octet?

a- P $\boxed{\text{b-C}}$ c-Cl d-Xe e-S



1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)

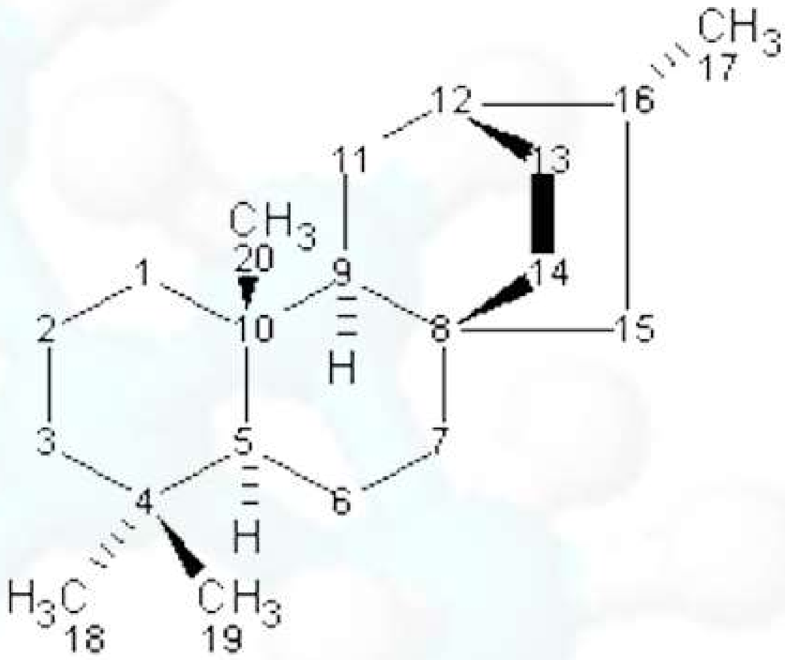
58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals

▨ Metalloids

🟦 Nonmetals

Thank you



Chapter Fourteen

Chemical Equilibrium

The Concept of Equilibrium and the Equilibrium Constant

- Few chemical reactions proceed in only one direction, most are reversible.
- At the start of a reversible process, the reaction proceeds toward the formation of products. As soon as some product molecules are formed, the reverse process begins to take place and reactant molecules are formed from product molecules.

$$A + B \leftrightarrow C + D$$

- Chemical equilibrium is achieved when:
- The rates of the forward and reverse reactions are equal and
- The concentrations of the reactants and products remain constant

Physical equilibrium

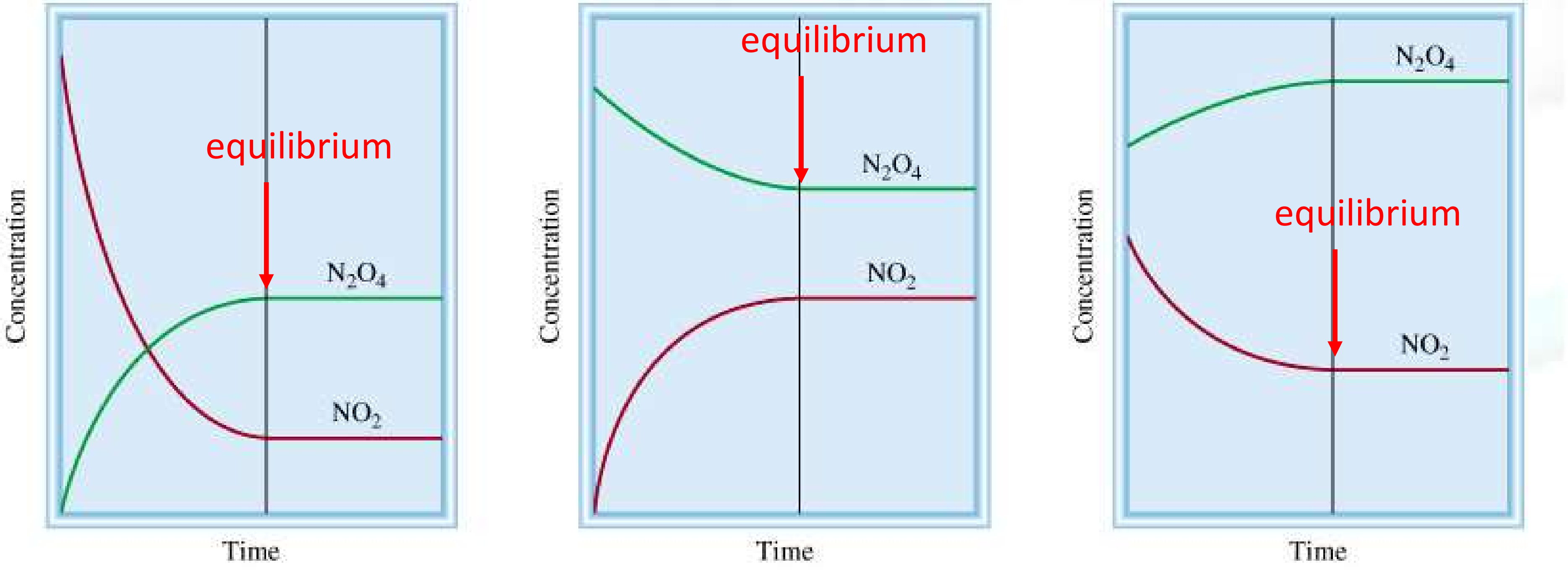
$$H_2O\ (l) \rightleftarrows H_2O\ (g)$$

Chemical equilibrium

$$N_2O_4\ (g) \rightleftarrows 2NO_2\ (g)$$

The Concept of Equilibrium and the Equilibrium Constant

$$N_2O_4\ (g) \rightleftharpoons 2NO_2\ (g)$$



Start with NO_2

Start with N_2O_4

Start with NO_2 & N_2O_4

The Concept of Equilibrium and the Equilibrium Constant

$$aA + bB \rightleftharpoons cC + dD$$

- Where a, b, c, & d: are the stoichiometric coefficients for A, B, C , & D.

$$\boxed{K = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

Law of Mass Action

Where K is equilibrium constant.

for a reversible reaction at equilibrium and a constant temperature, a certain ratio of reactant and product concentrations has a constant value, K, called the equilibrium constant.

Always

The concentration of **solids** and **pure liquids** and **solvent** are not included in the expression for the equilibrium constant.

K dos not have a unit

The Concept of Equilibrium and the Equilibrium Constant

- The magnitude of the equilibrium constant tells us whether an equilibrium reaction favors the products or reactants.
- If K is much greater than 1 (that is, K >> 1), the equilibrium will lie to the right and favors the products.
- if the equilibrium constant is much smaller than 1 (that is, K << 1), the equilibrium will lie to the left and favor the reactants.
- In this context, any number greater than 10 is considered to be much greater than 1, and any number less than 0.1 is much less than 1.

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

- homogeneous equilibrium applies to reactions in which all reacting species are in the same phase.

Example

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

K can be given as

$$K_c = \frac{[NO_2]^2}{[N_2O_4]}$$

$$K_p = \frac{P^2_{NO_2}}{P_{N_2O_4}}$$

Note that the subscript in K_c indicates that the concentrations of the reacting species are expressed in molarity or moles per liter.

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

Relationship between Kc and Kp:

Kc ≠ Kp

$$\boxed{K_p = K_c(RT)^{\Delta n}}$$

Δn = moles of gaseous products – moles of gaseous reactants

$$aA + bB \leftrightarrow cC + dD$$

$$\Delta n = (c+d)_{\text{products (g)}} - (a+b)_{\text{reactants (g)}}$$

Kc = Kp, when Δn = 0

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

Example

Write expressions for K_c , and K_p if applicable, for the following reversible reactions at equilibrium?

$$(a) \qquad HF(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + F^-(aq)$$

$$(b) \qquad 2NO(g) + O_2(g) \leftrightarrow 2NO_2(g)$$

$$(c) \qquad CH_3COOH(aq) + C_2H_5OH(aq) \leftrightarrow CH_3COOC_2H_5(aq) + H_2O(l)$$

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

(a) $$HF(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + F^-(aq)$$

$$\boxed{K = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

$$K_c = \frac{[H_3O^+][F^-]}{[HF][H_2O]}$$

Always do not write solvent in the expression of equilibrium constant.

Normally water is solvent

$$K_c = \frac{[H_3O^+][F^-]}{[HF]}$$

Because there are no gases then there is no K_p

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

(b) $$2NO(g)+O_2(g)\leftrightarrow 2NO_2(g)$$

$$K_c=\frac{[NO_2]^2}{[NO]^2[O_2]}$$

$$K_p=\frac{P_{NO_2}^2}{P_{NO}^2P_{O_2}}$$

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

$$(c) \qquad CH_3COOH(aq) + C_2H_5OH(aq) \leftrightarrow CH_3COOC_2H_5(aq) + H_2O(l)$$

$$K_c = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]}$$

Always do not write solvent in the expression of equilibrium constant.

Normally water is solvent

$$K_c = \frac{[CH_3COOC_2H_5]}{[CH_3COOH][C_2H_5OH]}$$

Because there are no gases then there is no K_p

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

- Example

The following equilibrium process has been studied at 230 °C:

$$2NO(g) + O_2(g) \leftrightarrow 2NO_2(g)$$

In one experiment, the concentrations of the reacting species at equilibrium are found to be [NO] = 0.0542 M, $[O_2]$ = 0.127 M, and $[NO_2]$ = 15.5 M. Calculate the equilibrium constant (K_c) of this reaction at this temperature.

$$K_c = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

$$K_c = \frac{(15.5)^2}{(0.0542)^2(0.127)} = 6.44 \text{ x } 10^5$$

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

- Example

The equilibrium constant K_p for the decomposition of

$$PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$$

Is found to be 1.05 at 250 °C . If the equilibrium partial pressures of PCl_5, and PCl_3 are 0.875 atm and 0.463 atm, respectively, what is the equilibrium partial pressure of Cl_2 at 250 °C.

$$K_p = \frac{P_{PCl_3} P_{Cl_2}}{P_{PCl_5}}$$

$$P_{Cl_2} = \frac{K_p P_{PCl_5}}{P_{PCl_3}} = \frac{(1.05)(0.875)}{(0.463)} = 1.98\,\text{atm}$$

Writing Equilibrium Constant Expressions

Homogeneous Equilibria

- Example

The equilibrium concentrations for the reaction between carbon monoxide and molecular chlorine to form $COCl_2$ (*g*) at 74^0C are [CO] = 0.012 *M*, $[Cl_2]$ = 0.054 *M*, and $[COCl_2]$ = 0.14 *M*. Calculate the equilibrium constants K_c and K_p.

$$CO\ (g) + Cl_2\ (g) \rightleftarrows COCl_2\ (g)$$

$$K_c = \frac{[COCl_2]}{[CO][Cl_2]}$$

$$= \frac{0.14}{0.012 \times 0.054} = 220$$

$$K_p = K_c(RT)^{\Delta n}$$

Δn = moles of products – moles of reactants

$\Delta n = 1 - 2 = -1$ $R = 0.0821$

$T = 273 + 74 = 347$ K

$K_p = 220 \times (0.0821 \times 347)^{-1} = 7.7$

Writing Equilibrium Constant Expressions

Heterogeneous Equilibria

- heterogeneous equilibrium applies to reactions in which all reacting species are in different phases.

Example

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

- All rules applied for homogeneous equilibria also applies for heterogeneous.

Writing Equilibrium Constant Expressions

Heterogeneous Equilibria

- Example

Write the equilibrium constant expression K_c, and K_p if applicable, for each of the following heterogeneous systems:

$$(a) \qquad (NH_4)_2Se(s) \leftrightarrow 2NH_3(g) + H_2Se(g)$$

$$(b) \qquad AgCl(s) \leftrightarrow Ag^+(aq) + Cl^-(aq)$$

$$(c) \qquad P_4(s) + 6Cl_2(g) \leftrightarrow 4PCl_3(l)$$

Writing Equilibrium Constant Expressions

Heterogeneous Equilibria

(a) $$(NH_4)_2Se(s) \leftrightarrow 2NH_3(g) + H_2Se(g)$$

$$K_c = \xcancel{\frac{[NH_3]^2[H_2Se]}{[(NH_4)_2Se]}}$$

$$K_c = [NH_3]^2[H_2Se]$$

$$K_p = P_{NH_3}^2 P_{H_2Se}$$

Writing Equilibrium Constant Expressions

Heterogeneous Equilibria

$$(b) \qquad AgCl(s) \leftrightarrow Ag^{+}(aq) + Cl^{-}(aq)$$

- AgCl is solid therefore it dose not count in equilibrium constant.

$$K_c = [Ag^{+}][Cl^{-}]$$

There are no gases thus there are no K_p.

Writing Equilibrium Constant Expressions

Heterogeneous Equilibria

$$(c) \qquad P_4(s) + 6Cl_2(g) \leftrightarrow 4PCl_3(l)$$

- P_4 is solid and PCl_3 is liquid therefore they do not count in constant of equilibrium.

$$K_c = \frac{1}{[Cl_2]^6}$$

$$K_p = \frac{1}{P_{Cl_2}^6}$$

Writing Equilibrium Constant Expressions

Heterogeneous Equilibria

- Example

Consider the following heterogeneous equilibrium:

$$CaCO_3(s) \leftrightarrow CaO(s) + CO_2(g)$$

At 800 °C, the pressure of CO_2 is 0.236 atm. Calculate (a) K_p and (b) K_c for the reaction at this temperature.

$CaCO_3$ and CaO are solid therefore they do not count in equilibrium of constant.

$$(a)\ \ K_p = P_{CO_2} = 0.236$$

$$K_p = K_c(RT)^{\Delta n}$$

$$K_c = \frac{K_p}{(RxT)^{\Delta n}}$$

(b) $R = 0.0821$

$T = 273 + 800 = 1073$ K

Δn = moles of products – moles of reactants

$\Delta n = 1 – 0 = 1$

$$K_c = \frac{0.236}{(0.0821x1073)^1} = 2.68x10^{-3}$$

Writing Equilibrium Constant Expressions

Multiple Equilibria

$$A + B \rightleftharpoons \cancel{C} + \cancel{D} \qquad K_c' \qquad K_c' = \frac{[C][D]}{[A][B]}$$

$$\cancel{C} + \cancel{D} \rightleftharpoons E + F \qquad K_c'' \qquad K_c'' = \frac{[E][F]}{[C][D]}$$

$$\overline{}$$

$$A + B \rightleftharpoons E + F \qquad K_c \qquad K_c = \frac{[E][F]}{[A][B]}$$

$$\boxed{K_c = K_c' \times K_c''}$$

- If a reaction can be expressed as the sum of two or more reactions, the equilibrium constant for the overall reaction is given by the product of the equilibrium constants of the individual reactions.

Writing Equilibrium Constant Expressions

Multiple Equilibria

- When the equation for a reversible reaction is written in the opposite direction, the equilibrium constant becomes the reciprocal of the original equilibrium constant. Thus, if we write the NO_2 –N_2O_4 equilibrium as

$$N_2O_4\ (g) \rightleftharpoons 2NO_2\ (g)$$

$$K = \frac{[NO_2]^2}{[N_2O_4]} = 4.63 \times 10^{-3}$$

- However, we can represent the equilibrium equally well as

$$2NO_2\ (g) \rightleftharpoons N_2O_4\ (g)$$

$$'K = \frac{[N_2O_4]}{[NO_2]^2} = \frac{1}{K} = 216$$

When the equation for a reversible reaction is written in the opposite direction, the equilibrium constant becomes the reciprocal of the original equilibrium constant

What Does the Equilibrium Constant Tell Us

Predicting the Direction of a Reaction

Example: The equilibrium constant for the following reaction at 700 K is,

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad K_1 = 10.17$$

What is the value of the equilibrium constant for the following :

$$HI(g) \rightleftharpoons 1/2\ H_2(g) + 1/2\ I_2(g) \qquad K_2 = ?$$

a- 9.668 x10^{-3} b- 3.189 c- 0.314 d- 0.098

$$K_2 = \sqrt{1/K_1}$$

$$K_2 = \sqrt{1/10.17} = 0.314$$

What Does the Equilibrium Constant Tell Us

Predicting the Direction of a Reaction

Example: The equilibrium constant for the following reaction at 700 K is,

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad K_1 = 10.17$$

What is the value of the equilibrium constant for the following :

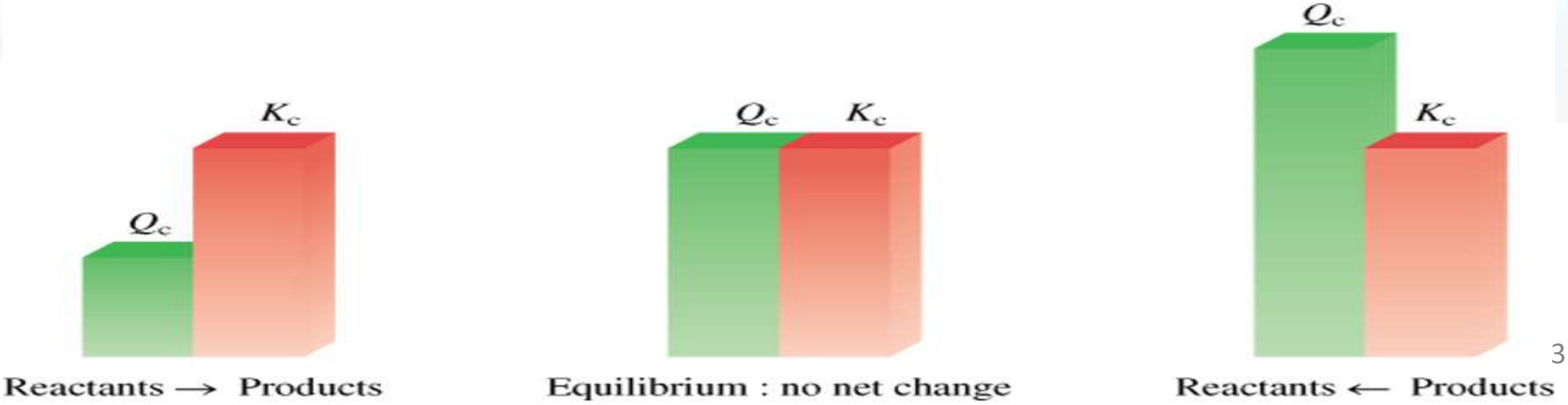
$$HI(g) \rightleftharpoons 1/2\ H_2(g) + 1/2\ I_2(g) \qquad K_2 = ?$$

a- 9.668×10^{-3} b- 3.189 [✓ c- 0.314] d- 0.098

What Does the Equilibrium Constant Tell Us

Predicting the Direction of a Reaction

- reaction quotient (Q_c): is calculated by substituting the initial concentrations of the reactants and products into the equilibrium constant (K_c) expression.
- **IF**
- $Q_c > K_c$ system proceeds from right to left to reach equilibrium
- $Q_c = K_c$ the system is at equilibrium
- $Q_c < K_c$ system proceeds from left to right to reach equilibrium



What Does the Equilibrium Constant Tell Us

Predicting the Direction of a Reaction

- The equilibrium constant K_c for the formation of hydrogen iodide from molecular hydrogen and molecular iodine in the gas phase

$$H_2(g) + I_2(g) \leftrightarrow 2HI(g)$$

is 54.3 at 430°C. that in a certain experiment we place 0.243 mole of H_2, 0.146 mole of I_2, and 1.98 moles of HI all in a 1.00-L container at 430°C.

$$\frac{[HI]_o^2}{[H_2]_o[I_2]_o} = \frac{(1.98)^2}{(0.243)(0.146)} = 111$$

where the subscript 0 indicates initial concentrations (before equilibrium is reached). Because the quotient $[HI]_0^2 / [H_2]_0 [I_2]_0$ is greater than K_c, this system is not at equilibrium

What Does the Equilibrium Constant Tell Us

Predicting the Direction of a Reaction

- Example:

At the start of a reaction, there are 0.249 mol N_2, 3.21 x 10^{-2} mol H_2, and 6.42 x 10^{-4} mol NH_3 in a 3.50 L reaction vessel at 375 °C. If the equilibrium constant (K_c) for the reaction:

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$$

Is 1.2 at this temperature, decide whether the system is at equilibrium. If it is not, predict which way the net reaction proceed?

$$[N_2]_o = \frac{0.249\,\text{mol}}{3.50\,\text{L}} = 0.0711\,\text{M}$$

$$[H_2]_o = \frac{3.21x10^{-2}\ \text{mol}}{3.50\,\text{L}} = 9.17x10^{-3}\ \text{M}$$

$$[NH_3]_o = \frac{6.42x10^{-4}\ \text{mol}}{3.50\,\text{L}} = 1.83x10^{-4}\ \text{M}$$

$$Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(1.83x10^{-4})^2}{(0.0711)(9.17x10^{-3})^2} = 0.611$$

$Q_c < K_c$

THUS: the reaction is NOT at equilibrium

THUS: the product concentration will increase and the reactant concentration will decrease

The reaction will proceed from left to right until the equilibrium is reached.

What Does the Equilibrium Constant Tell Us

Calculating Equilibrium Concentrations

- If we know the equilibrium constant for a particular reaction, we can calculate the concentrations in the equilibrium mixture from the initial concentrations.

1. Express the equilibrium concentrations of all species in terms of the initial concentrations and a single unknown x, which represents the change in concentration.
2. Write the equilibrium constant expression in terms of the equilibrium concentrations. Knowing the value of the equilibrium constant, solve for x.
3. Having solved for x, calculate the equilibrium concentrations of all species.

What Does the Equilibrium Constant Tell Us

Calculating Equilibrium Concentrations

- Example

A mixture of 0.500 mol H_2 and 0.500 mol I_2 was placed in a 1.00 L stainless-steel flask at 430 °C. The equilibrium constant K_c for the reaction

$$H_2(g) + I_2(g) \leftrightarrow 2HI(g)$$

Is 54.3 at this temperature. Calculate the concentrations of H_2, I_2 and HI at equilibrium

	H_2	I_2	2HI
Initial (M)	0.5	0.5	0.0
Change	-x	-x	+2x
Equilibrium	0.5 - x	0.5 - x	2x

What Does the Equilibrium Constant Tell Us

Calculating Equilibrium Concentrations

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$54.3 = \frac{[2x]^2}{[0.5-x][0.5-x]} = \frac{[2x]^2}{[0.5-x]^2}$$

$$\sqrt{54.3} = \frac{\sqrt{[2x]^2}}{\sqrt{[0.5-x]^2}}$$

$$7.369 = \frac{2x}{0.5-x}$$

$$7.369(0.5-x) = 2x$$

$$3.684 - 7.369x = 2x$$

$$3.684 = 2x + 7.369x$$

$$3.684 = 9.369x$$

$$x = \frac{3.684}{9.369} = 0.393$$

$$[H_2] = (0.5-0.393) = 0.107\text{ M}$$

$$[I_2] = (0.5-0.393) = 0.107\text{ M}$$

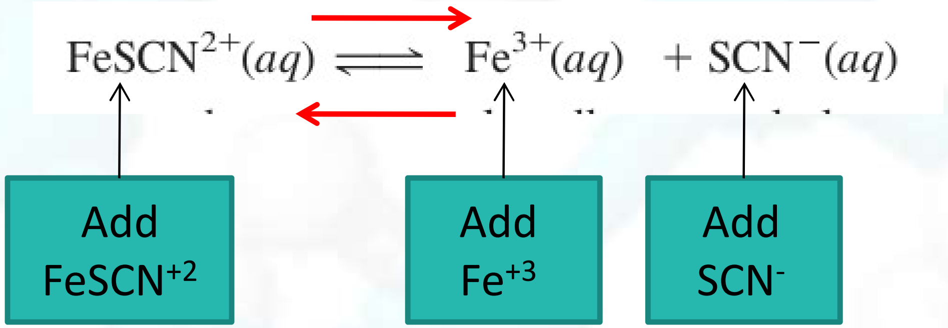
$$[HI] = 2(0.393) = 0.786\text{ M}$$

Factors That Affect Chemical Equilibrium

- Chemical equilibrium represents a balance between forward and reverse reactions. In most cases, this balance is quite delicate. Changes in experimental conditions may disturb the balance and shift the equilibrium position so that more or less of the desired product is formed.
- In this section we will study 5 factor which can effect chemical equilibrium namely : concentration, pressure, volume, temperature, and catalyst.
- **Le Châtelier's Principle**: if an external stress is applied to a system at equilibrium, the system adjusts in such a way that the stress is partially offset as the system reaches a new equilibrium position.

Factors That Affect Chemical Equilibrium

Changes in Concentration



$$FeSCN^{2+}(aq) \rightleftharpoons Fe^{3+}(aq) + SCN^{-}(aq)$$

Add $FeSCN^{+2}$

Add Fe^{+3}

Add SCN^{-}

Change	**Shifts the Equilibrium**
Increase concentration of product(s)	Left (toward reactant)
Decrease concentration of product(s)	Right (toward product)
Increase concentration of reactant(s)	Right (toward product)
Decrease concentration of reactant(s)	Left (toward reactant)

Concentration effect the **position** and **not the value** of K

Factors That Affect Chemical Equilibrium

Changes in Concentration

Example

If all species are gases and H_2 is added, the amount of CO present at equilibrium will:

$$CO_2 + H_2 \leftrightarrow CO + H_2O$$

[✓ a- Increase.] b-decrease. c-remain unchanged. d-disappear.

Factors That Affect Chemical Equilibrium

Changes in Concentration

Example

At 720 °C, the equilibrium K_c for the reaction:

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$$

Is 2.37 x 10^{-3}. the equilibrium concentrations are $[N_2]$ = 0.683 M,$[H_2]$ = 8.80 M, and $[NH_3]$ = 1.05 M. Suppose some NH_3 is added to the mixture so that the concentration is increased to 3.65 M. (a) Use Le Châtelier's Principal to predict the shift direction of the net reaction to reach new equilibrium. (b) Confirm your prediction by calculating the reaction quotient Q_c and comparing its value with K_c.

Factors That Affect Chemical Equilibrium

Changes in Concentration

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$$

a-The increase was in product thus the equilibrium will shift toward reactant (left).

b-

$$Q_c = \frac{[NH_3]_o^2}{[N_2]_o[H_2]_o^3}$$

$$Q_c = \frac{(3.65)^2}{(0.683)(8.80)^3} = 2.86x10^{-2}$$

$Q_c > K_c$

$Q_c > K_c$ system proceeds from right to left to reach equilibrium ($Q_c = K_c$)

Factors That Affect Chemical Equilibrium

Changes in Volume and Pressure

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

- In general, an increase in pressure (decrease in volume) favors the net reaction that decreases the total number of moles of gases.
- a decrease in pressure (increase in volume) favors the net reaction that increases the total number of moles of gases
- For reactions in which there is no change in the number of moles of gases, a pressure (or volume) change has no effect on the position of equilibrium.
- Volume has the opposite effect of pressure

Factors That Affect Chemical Equilibrium

Changes in Volume and Pressure

Example

Consider the following equilibrium systems:

$$(a)\ 2PbS(s)+3O_2(g)\leftrightarrow 2PbO(s)+2SO_2(g)$$

$$(b)\ PCl_5(g)\leftrightarrow PCl_3(g)+Cl_2(g)$$

$$(c)\ H_2(g)+CO_2(g)\leftrightarrow H_2O(g)+CO(g)$$

Predict the direction of the net reaction in each case as a result of increasing the pressure (decreasing the volume) on the system at constant temperature.?

Factors That Affect Chemical Equilibrium

Changes in Volume and Pressure

$$(a)\ \ 2PbS(s) + 3O_2(g) \leftrightarrow 2PbO(s) + 2SO_2(g)$$

- 3 mole for reactant and 2 mole for product

Thus it will go toward product (less mole).

$$(b)\ \ PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$$

- 1 mole for reactant and 2 mole for product

Thus it will go toward reactant (less mole)

$$(c)\ \ H_2(g) + CO_2(g) \leftrightarrow H_2O(g) + CO(g)$$

- 2 mole for reactant and 2 mole for product

Thus it will remain unchanged

Factors That Affect Chemical Equilibrium

Changes in Temperature

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

- The forward reaction is endothermic (absorbs heat, ΔH° > 0):

$$\text{heat} + N_2O_4(g) \longrightarrow 2NO_2(g) \qquad \Delta H^\circ = 58.0\ \text{kJ/mol}$$

- The reverse reaction is exothermic (releases heat, ΔH° < 0)

$$2NO_2(g) \longrightarrow N_2O_4(g) + \text{heat} \qquad \Delta H^\circ = -58.0\ \text{kJ/mol}$$

- Temperature increase favour the endothermic reaction,
- Temperature decrease favours an exothermic reaction
- Only a change in temperature can alter the equilibrium constant.

Factors That Affect Chemical Equilibrium

Changes in Temperature

Example

If the reaction is endothermic and the temperature is raised, the amount of CO present will:

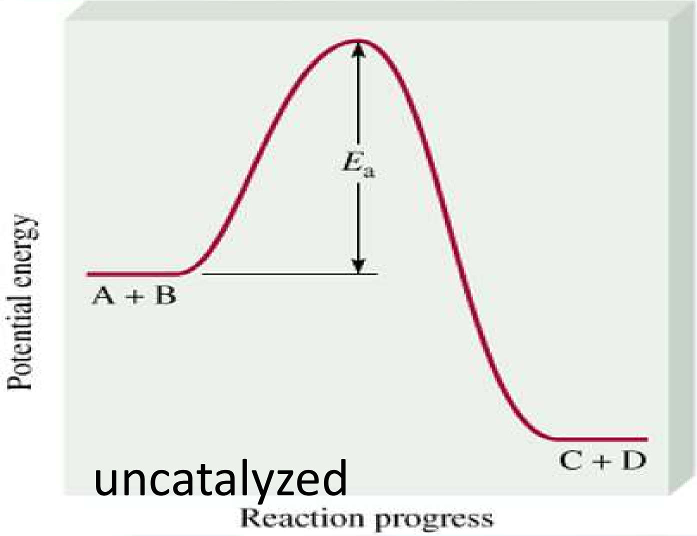
$$CO_2 + H_2 \leftrightarrow CO + H_2O$$

[✓ a-increase.] b-decrease. c-remain unchanged. d-disappear.

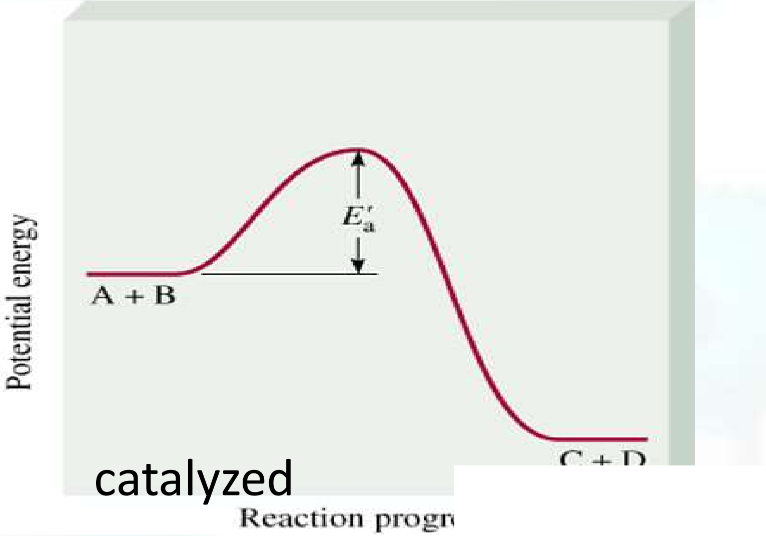
Factors That Affect Chemical Equilibrium

The Effect of a Catalyst

- Adding a Catalyst
  - does not change *K*
  - does not shift the position of an equilibrium system
  - system will reach equilibrium sooner



uncatalyzed



catalyzed

- Catalyst lowers E_a for **both** forward and reverse reactions.
- Catalyst does not change equilibrium constant or shift equilibrium.

Factors That Affect Chemical Equilibrium

The Effect of a Catalyst

Example

If a catalyst is added to the equilibrium, the amount of CO present will:

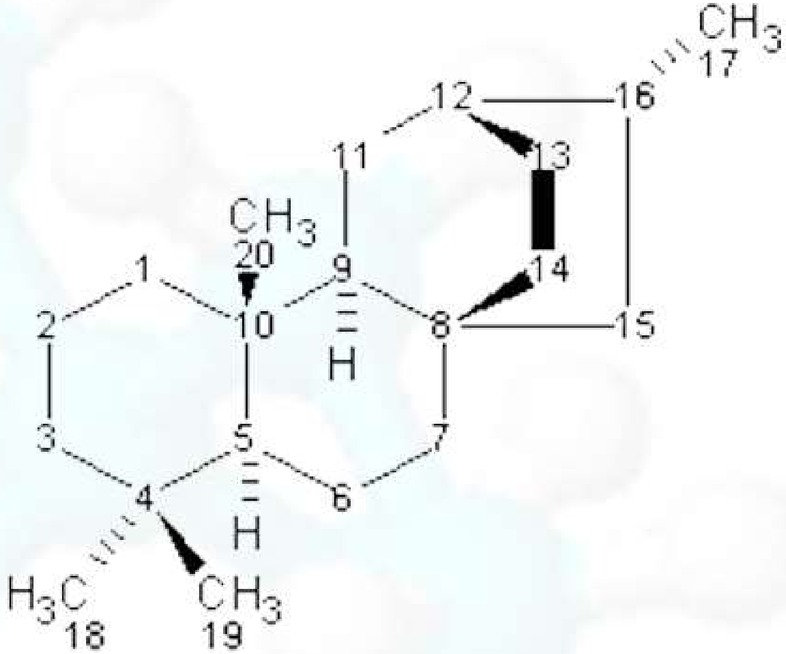
$$CO_2 + H_2 \leftrightarrow CO + H_2O$$

a-increase. b-decrease. [✓ c-remain unchanged.] d-disappear.

Factors That Affect Chemical Equilibrium

Change	Shift Equilibrium	Change Equilibrium Constant
Concentration	yes	no
Pressure	yes	no
Volume	yes	no
Temperature	yes	yes
Catalyst	no	no

Thank you



Chapter Fifteen

Acids and Bases

The Acid-Base Properties of Water

- acids is a substances that ionize in water to produce H^+ ions

$$HCl\ (aq) \rightarrow H^+(aq) + Cl^-(aq)$$

- bases is a substances that ionize in water to produce OH^- ions.

$$NaOH\ (aq) \rightarrow Na^+(aq) + OH^-(aq)$$

- An acid neutralizes a base

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$$

The Acid-Base Properties of Water

- Water has the ability to act either as an acid or as a base.
- Water undergo ionization to a small extent this reaction is sometimes called the autoionization of water.

$$H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq)$$

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

$$K_c = [H_3O^+][OH^-] = [H^+][OH^-]$$

$$K_W = [H^+][OH^-]$$

- The ***ion-product constant*** (***K_w***) is the product of the molar concentrations of H^+ and OH^- ions **at a particular temperature.**
- **At 25 °C, K_w = 1.0 × 10^{-14}**

The Acid-Base Properties of Water

For water:

$$K_w = [H_3O^+][HO^-] = [H^+][HO^-] = 1x10^{-14}$$

Because water is neutral then

$$[H^+] = [HO^-] = \sqrt{1x10^{-14}} = 1x10^{-7}\ \mathrm{M}$$

	Solution Is
$[H^+] = [OH^-]$	neutral
$[H^+] > [OH^-]$	acidic
$[H^+] < [OH^-]$	basic

The Acid-Base Properties of Water

Example:

Calculate the $[H^+]$ ions in aqueous ammonia , $[OH^-]$ =0.0025 M?

$$K_W = [H^+][OH^-]$$

$$[H^+] = \frac{K_W}{[OH^-]}$$

$$[H^+] = \frac{1x10^{-14}}{0.0025} = 4x10^{-12} M$$

THUS $[H^+] < [OH^-]$

therefore the solution is basic

pH—A Measure of Acidity

- Because the concentrations of H^+ and OH^- ions in aqueous solutions are frequently very small numbers and therefore inconvenient to work with, Soren Sorensen in 1909 proposed a more practical measure called pH.
- The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration (in mol/L).

$$pH = -\log[H^+] = -\log[H_3O^+]$$

$$[H^+] = 10^{-pH}$$

- For [OH]

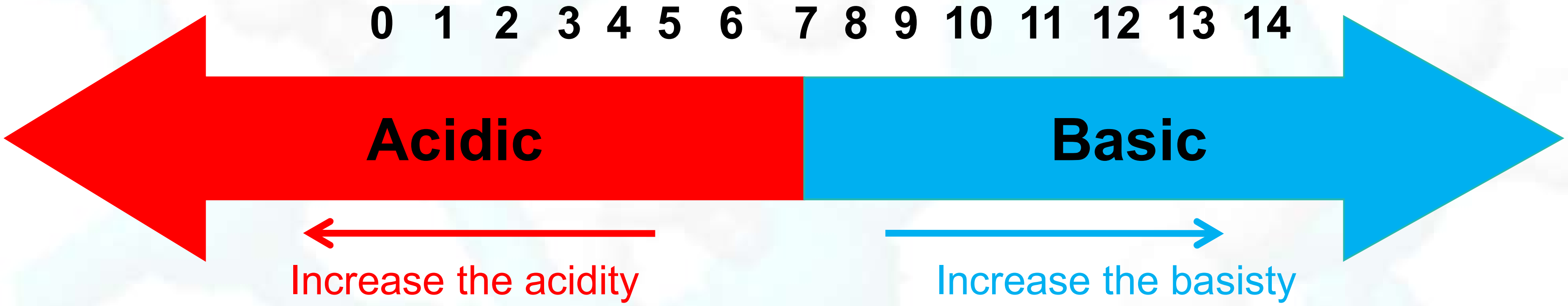
$$pOH = -\log[OH^-]$$

$$[OH^-] = 10^{-pOH}$$

$$pK_w = -\log 1x10^{-14} = 14$$

$$pH + pOH = 14$$

pH—A Measure of Acidity



Solution Is		**At 25^0C**	
neutral	$[H^+] = [OH^-]$	$[H^+] = 1 \times 10^{-7}$	pH = 7
acidic	$[H^+] > [OH^-]$	$[H^+] > 1 \times 10^{-7}$	pH < 7
basic	$[H^+] < [OH^-]$	$[H^+] < 1 \times 10^{-7}$	pH > 7

pH is inversely proportional to $[H^+]$

pH—A Measure of Acidity

Example

The concentration of H^+ ions in a bottle of vinegar was 3.2 x 10^{-4} M right after the cork was removed. Only half of the vinegar was consumed. The other half, after it had been standing open to the air for a month, was found to have a hydrogen ion concentration equal to 1.0 x 10^{-3} M. Calculate the pH of the vinegar on these two occasions.

$$pH = -\log[H^+] = -\log(3.2x10^{-4}) = 3.49$$

$$pH = -\log[H^+] = -\log(1.0x10^{-3}) = 3.00$$

pH—A Measure of Acidity

Example

The pH of rainwater collected in a certain region of Saudi Arabia on a particular day was 4.82. Calculate the H^+ ion concentration of the rainwater.

$$pH = -\log[H^+]$$

$$[H^+] = 10^{-pH}$$

$$[H^+] = 10^{-4.82} = 1.5x10^{-5} M$$

pH—A Measure of Acidity

Example

In a NaOH solution $[OH^-]$ is 2.9×10^{-4} M. Calculate the pH of the solution?

$$pOH = -\log[OH^-] = -\log(2.9x10^{-4}) = 3.54$$

$$pH + pOH = 14$$

$$pH = 14 - pOH = 14 - 3.54 = 10.46$$

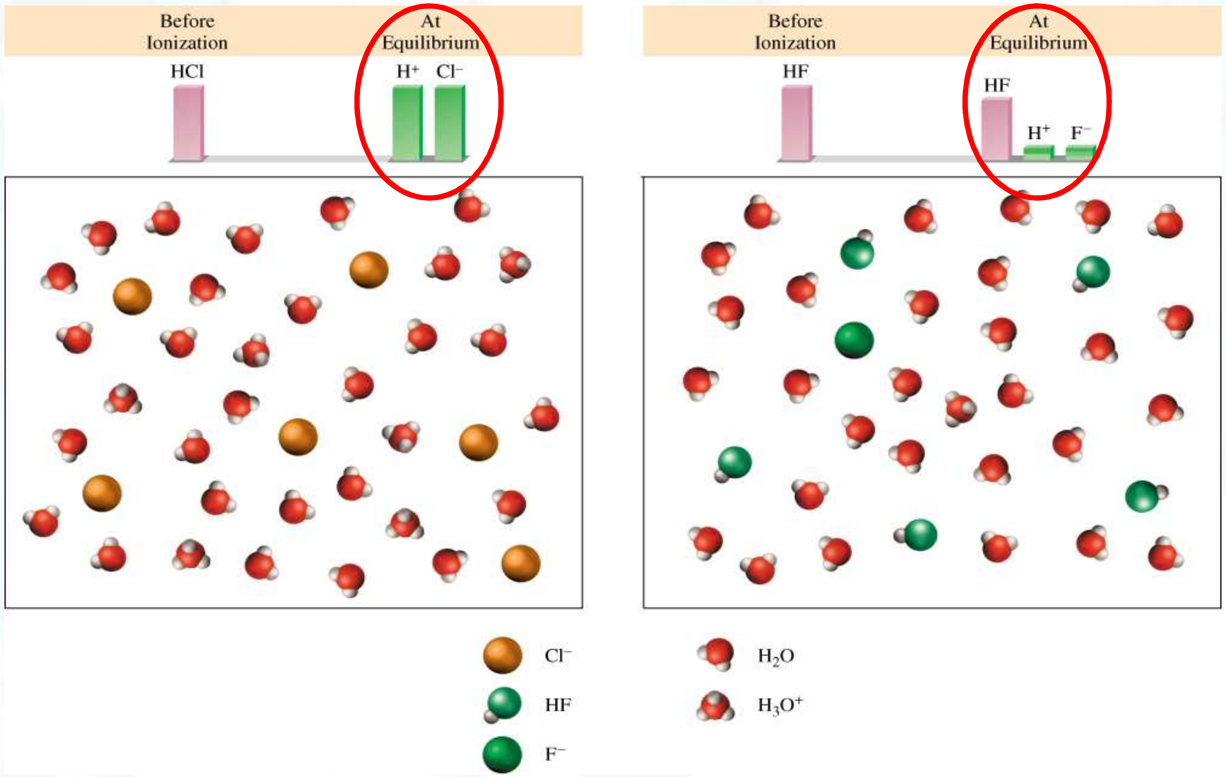
Weak Acids and Acid Ionization Constants

- Strong acid (or base) have 100 % dissociation.

$$HCl\ (s) \xrightarrow{H_2O} H^+\ (aq) + Cl^-\ (aq)$$

- Weak acid (or base) have incomplete dissociation.

$$CH_3COOH \rightleftarrows CH_3COO^-\ (aq) + H^+\ (aq)$$



Weak Acids and Acid Ionization Constants

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

- The acid ionization constant (K_a), is the equilibrium constant for the ionization of an acid.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

- At a given temperature, the strength of the acid HA is measured quantitatively by the magnitude of K_a. The larger K_a, the stronger the acid that is, the grater the concentration of H^+ ions at equilibrium due to its ionization.

$$[H^+] = \sqrt{K_a[acid]}$$

$$[OH^-] = \sqrt{K_b[base]}$$

Weak Acids and Acid Ionization Constants

Example

What is the pH of a 0.5 *M* HF solution (at 25^0C) if K_a= $7.1x10^{-4}$?

$$HF(aq) \rightleftharpoons H^{+}(aq) + F^{-}(aq)$$

$$[H^+] = \sqrt{K_a[acid]}$$

$$[H^+] = \sqrt{7.1x10^{-4}x0.5}$$

$$[H^+] = 0.019$$

pH = -log [H^+] = 1.72

Weak Acids and Acid Ionization Constants

Example

- What is the pH of a 0.122 *M* monoprotic acid whose K_a is 5.7 x 10^{-4}?

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

$$[H^+] = \sqrt{K_a[acid]}$$

$$[H^+] = \sqrt{5.7x10^{-4}x0.122}$$

$$[H^+] = 0.008$$

pH = -log $[H^+]$ = 2.08

Weak Acids and Acid Ionization Constants

Example

The pH of a 0.10 M solution of formic acid (HCOOH) is 2.39. What is the K_a of the acid?

$$HCOOH\ (aq) \rightleftharpoons H^+\ (aq) + HCOO^-\ (aq)$$

$$pH = -\log[H^+]$$

$$[H^+] = 10^{-pH}$$

$$[H^+] = 10^{-2.39} = 4.1x10^{-3}M$$

$$[H^+] = \sqrt{K_a[acid]}$$

$$[H^+]^2 = K_a[acid]$$

$$K_a = \frac{[H^+]^2}{[acid]}$$

$$K_a = \frac{[4.1x10^{-3}]^2}{[0.1]}$$

$$K_a = 1.7x10^{-4}$$

Weak Acids and Acid Ionization Constants

- K_a indicates the strength of an acid. Another measure of the strength of an acid is percent ionization.

$$\textbf{\textit{percent ionization}} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\%$$

$$\text{Percent ionization} = \frac{[H^+]}{[HA]_0} \times 100\%$$

$[HA]_0$ = initial concentration

- The stronger the acid, the greater the percent ionization.

Weak Acids and Acid Ionization Constants

Example

Calculate the percent ionization of hydrofluoric acid at the concentrations of 0.50 M if K_a= 7.1x10^{-4} ?

$$\text{HF }(aq) \rightleftharpoons \text{H}^+ (aq) + \text{F}^- (aq)$$

$$[H^+] = \sqrt{K_a x[acid]}$$

$$[H^+] = \sqrt{7.1x10^{-4} x0.5}$$

$$[H^+] = 0.019M$$

$$\textbf{\textit{percent ionization}} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \text{ x } 100\%$$

$$\textbf{\textit{percent ionization}} = \frac{0.019}{0.5} \text{ x } 100\% = 3.8\%$$

Weak Acids and Acid Ionization Constants

Example

A 0.040 M solution of a monoprotic acid is 3 percent ionized. Calculate the ionization constant of the acid.?

$$\textbf{\textit{percent ionization}} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\%$$

$$3 = \frac{[H^+]}{0.04} X100$$

$$[H^+] = \frac{0.04x3}{100} = 0.0012M$$

$$[H^+] = \sqrt{K_a[acid]}$$

$$[H^+]^2 = K_a[acid]$$

$$K_a = \frac{[H^+]^2}{[acid]}$$

$$K_a = \frac{[0.0012]^2}{[0.04]}$$

$$K_a = 3.6x10^{-5}$$

Weak Bases

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

- The ionization of weak bases is treated in the same way as the ionization of weak acids.
- The base ionization constant (K_b), is the equilibrium constant for the ionization of a base.

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

- At a given temperature, the strength of the base BA is measured quantitatively by the magnitude of K_b. The larger K_b, the stronger the base—that is, the greater the concentration of OH^- ions at equilibrium due to its ionization
- In solving problems involving weak bases, we follow the same procedure we used for weak acids. The main difference is that we calculate $[OH^-]$ first, rather than $[H^+]$.

Weak Bases

- Example

What is the pH of a 0.40 M ammonia solution if K_b= 1.8x10^{-5}?

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$[OH^-] = \sqrt{K_b[base]}$$

$$[OH^-] = \sqrt{1.8x10^{-5}x0.4}$$

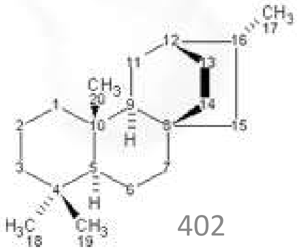
$$[OH^-] = 0.0027$$

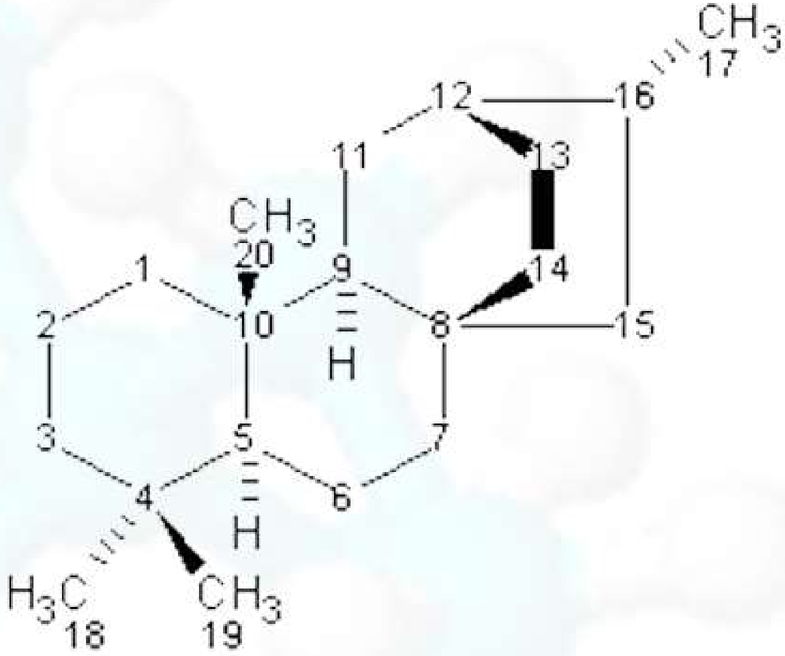
pOH = -log [OH$^-$] = 2.57

$$pH + pOH = 14$$

$$pH = 14 - pOH = 14 - 2.57 = 11.43$$

Thank you





Chapter Sixteen

Acid-Base Equilibria and Solubility Equilibria

The Common Ion Effect

$$CH_3COONa\ (s) \longrightarrow Na^+\ (aq)\ \boxed{+\ CH_3COO^-\ (aq)}$$

common ion

$$CH_3COOH\ (aq) \rightleftharpoons H^+\ (aq)\ \boxed{+\ CH_3COO^-\ (aq)}$$

$$\longleftarrow$$

- The ***common ion effect*** is the shift in equilibrium caused by the addition of a compound having an ion in common with the dissolved substance.

The Common Ion Effect

- Consider mixture of salt NaA and weak acid HA.

$$\mathrm{NaA}\ (s) \longrightarrow \mathrm{Na}^+\ (aq) + \mathrm{A}^-\ (aq)$$

$$\mathrm{HA}\ (aq) \rightleftharpoons \mathrm{H}^+\ (aq) + \mathrm{A}^-\ (aq)$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$[\mathrm{H}^+] = \frac{K_a\,[\mathrm{HA}]}{[\mathrm{A}^-]}$$

$$-\log\,[\mathrm{H}^+] = -\log K_a - \log \frac{[\mathrm{HA}]}{[\mathrm{A}^-]}$$

$$-\log\,[\mathrm{H}^+] = -\log K_a + \log \frac{[\mathrm{A}^-]}{[\mathrm{HA}]}$$

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

salt

Weak acid

$\mathrm{p}K_a = -\log K_a$

Henderson-Hasselbalch equation

The Common Ion Effect

Example

What is the pH of a solution containing both 0.20 M CH_3COOH and 0.30 M CH_3COONa? The K_a of CH_3COOH is 1.8×10^{-5}.

$$CH_3COOH\ (aq) \rightleftharpoons H^+\ (aq) + CH_3COO^-\ (aq)$$

$$CH_3COONa\ (s) \longrightarrow Na^+\ (aq) + CH_3COO^-\ (aq)$$

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

$$pH = -\log 1.8x10^{-5} + \log\frac{0.3}{0.2}$$

pH= 4.92

Buffer Solutions

- A buffer solution is a solution of (1) a weak acid or a weak base and (2) its salt; both components must be present.
- The solution has the ability to resist changes in pH upon the addition of small amounts of either acid or base.
- A buffer solution must contain a relatively large concentration of acid to react with any OH^- ions that are added to it, and it must contain a similar concentration of base to react with any added H^+ ions. Furthermore, the acid and the base components of the buffer must not consume each other in a neutralization reaction. These requirements are satisfied by an acid-base conjugate pair.

Buffer Solutions

- A solution containing acetic acid (CH_3COOH) and its salt sodium acetate (CH_3COONa) added to water these two substances has the ability to neutralize either added acid or added base. Sodium acetate, a strong electrolyte, dissociates completely in water:

$$CH_3COONa \longrightarrow CH_3COO^- (aq) + Na^+ (aq)$$

- Acetic acid is weak acid :

$$CH_3COOH (aq) \rightleftharpoons H^+ (aq) + CH_3COO^- (aq)$$

- If an acid is added, the H^+ ions will be consumed by the conjugate base in the buffer, CH_3COO^-, according to the equation:

$$H^+ (aq) + CH_3COO^- (aq) \longrightarrow CH_3COOH (aq)$$

- If a base is added to the buffer system, the OH^- ions will be neutralized by the acid in the buffer:

$$OH^- (aq) + CH_3COOH (aq) \longrightarrow CH_3COO^- (aq) + H_2O (l)$$

Buffer Solutions

Example

Which of the following are buffer systems? (a) KF/HF (b) KBr/HBr

(a) HF is a weak acid and F^- is its conjugate base

buffer solution

(b) HBr is a strong acid

not a buffer solution

Buffer Solutions

Example

A-What is the pH of a solution containing both 0.20 M CH_3COOH and 0.30 M CH_3COONa? The K_a of CH_3COOH is 1.8 x 10^{-5}.

B-If you add 0.01M HCl to the previous solution (buffer solution) what is the new PH?

C- If you add 0.01M NaOH to the previous solution (buffer solution) what is the new PH?

$$CH_3COOH\ (aq) \rightleftharpoons H^+\ (aq) + CH_3COO^-\ (aq)$$

Buffer Solutions

Example

A-What is the pH of a solution containing both 0.20 M CH_3COOH and 0.30 M CH_3COONa? The K_a of CH_3COOH is 1.8 x 10^{-5}.

$$CH_3COOH\ (aq) \rightleftharpoons H^+\ (aq) + CH_3COO^-\ (aq)$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = -\log 1.8 \times 10^{-5} + \log \frac{0.3}{0.2}$$

$$pH = 4.92$$

Buffer Solutions

Example

B-If you add 0.01M HCl to the previous solution (buffer solution) what is the new PH?

$$CH_3COOH\ (aq) \rightleftharpoons H^+\ (aq) + CH_3COO^-\ (aq)$$

HCl is strong acid therefore it equation will be as

$$HCl\ (aq) \rightarrow H^+(aq) + Cl^-(aq$$

Thus the H^+ will increase and the equilibrium will shift toward reactant.

Thus we have to add 0.01 to the weak acid concentration (CH_3COOH) so it will be (0.2+0.01=0.21)and also take 0.01 from the negative ion concentration (CH_3COO^-).(0.3-0.01= 0.29)

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = -\log 1.8 \times 10^{-5} + \log \frac{0.29}{0.21}$$

$$pH = 4.88$$

Buffer Solutions

Example

C- If you add 0.01M NaOH to the previous solution (buffer solution) what is the new PH?

$$CH_3COOH\ (aq) \rightleftharpoons H^+\ (aq) + CH_3COO^-\ (aq)$$

NaOH is strong base therefore it equation will be as

$$NaOH\ (aq) \rightarrow OH^-(aq) + Na^+(aq)$$

The OH^- ion produced from NaOH will react with H^+ ion. Then the H^+ will decrease and the equilibrium will shift toward product

Then we have to take0.01 to the weak acid concentration (CH_3COOH) so it will be (0.2-0.01=0.19)and also add 0.01 to the negative ion concentration (CH_3COO^-).(0.3+0.01= 0.31)

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = -\log 1.8 \times 10^{-5} + \log \frac{0.31}{0.19}$$

$$pH = 4.96$$

Preparing a Buffer Solution with a Specific pH

- to prepare a buffer solution, we work backwards. First we choose a weak acid whose pK_a is close to the desired pH. Next, we substitute the pH and pK_a values in Henderson-Hasselbalch equation to obtain the ratio [conjugate base]/[acid]. This ratio can then be converted to molar quantities for the preparation of the buffer solution.

- Smaller the pKa value, the stronger the acid.

Preparing a Buffer Solution with a Specific pH

Example:

Describe how you would prepare a “phosphate buffer” with a pH of about 7.40.?

$$H_3PO_4(aq) \rightleftharpoons H^+(aq) + H_2PO_4^-(aq) \quad K_{a_1} = 7.5 \times 10^{-3};\ pK_{a_1} = 2.12$$
$$H_2PO_4^-(aq) \rightleftharpoons H^+(aq) + HPO_4^{2-}(aq) \quad K_{a_2} = 6.2 \times 10^{-8};\ pK_{a_2} = 7.21$$
$$HPO_4^{2-}(aq) \rightleftharpoons H^+(aq) + PO_4^{3-}(aq) \quad K_{a_3} = 4.8 \times 10^{-13};\ pK_{a_3} = 12.32$$

The most suitable of the three buffer systems is HPO_4^{-2} /$H_2PO_4^-$, because the pK_a of the acid $H_2PO_4^-$ is closest to the desired pH. From the Henderson-Hasselbalch equation we write

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

$$7.4 = 7.21 + \log\frac{[A^-]}{[HA]}$$

$$0.19 = \log\frac{[A^-]}{[HA]}$$

$$10^{0.19} = \frac{[A^-]}{[HA]}$$

$$1.5 = \frac{[A^-]}{[HA]}$$

Preparing a Buffer Solution with a Specific pH

- Thus, one way to prepare a phosphate buffer with a pH of 7.40 is to dissolve disodium hydrogen phosphate (Na_2HPO_4) and sodium dihydrogen phosphate (NaH_2PO_4) in a mole ratio of 1.5:1.0 in water. For example, we could dissolve 1.5 moles of Na_2HPO_4 and 1.0 mole of NaH_2PO_4 in enough water to make up a 1-L solution

Solubility Equilibria

$$\mathrm{AgCl}\ (s) \rightleftharpoons \mathrm{Ag^+}\ (aq) + \mathrm{Cl^-}\ (aq)$$

$$K_{sp} = [\mathrm{Ag^+}][\mathrm{Cl^-}]$$

- The solubility product (K_{sp}) of a compound is the product of the molar concentrations of the constituent ions, each raised to the power of its stoichiometric coefficient in the equilibrium equation.

$\mathrm{MgF_2}\ (s) \rightleftharpoons \mathrm{Mg^{2+}}\ (aq) + 2\mathrm{F^-}\ (aq)$	$K_{sp} = [\mathrm{Mg^{2+}}][\mathrm{F^-}]^2$
$\mathrm{Ag_2CO_3}\ (s) \rightleftharpoons 2\mathrm{Ag^+}\ (aq) + \mathrm{CO_3^{2-}}\ (aq)$	$K_{sp} = [\mathrm{Ag^+}]^2[\mathrm{CO_3^{2-}}]$
$\mathrm{Ca_3(PO_4)_2}\ (s) \rightleftharpoons 3\mathrm{Ca^{2+}}\ (aq) + 2\mathrm{PO_4^{3-}}\ (aq)$	$K_{sp} = [\mathrm{Ca^{2+}}]^3[\mathrm{PO_4^{3-}}]^2$

- The value of K_{sp} indicates the solubility of an ionic compound, the smaller the value, the less soluble the compound in water.

Solubility Equilibria

- For concentrations of ions that do not correspond to equilibrium conditions we use the reaction quotient, which in this case is called the ion product (Q), to predict whether a precipitate will form.

Dissolution of an ionic solid in aqueous solution:

$Q < K_{sp}$	Unsaturated solution	No precipitate
$Q = K_{sp}$	Saturated solution	
$Q > K_{sp}$	Supersaturated solution	Precipitate will form

Solubility Equilibria

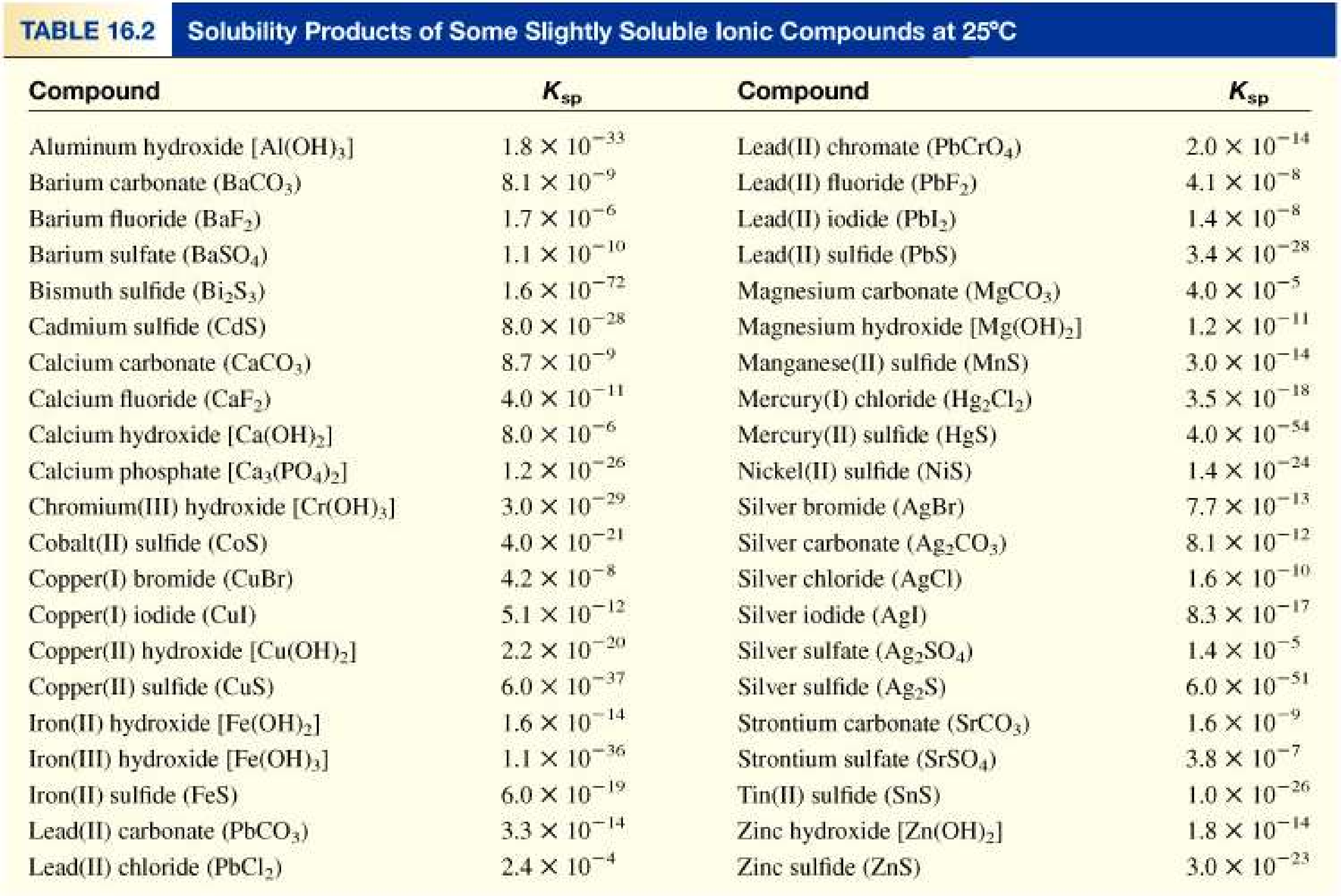
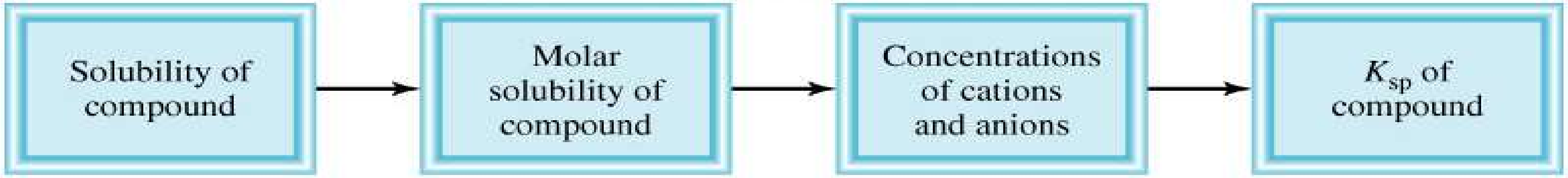


TABLE 16.2 **Solubility Products of Some Slightly Soluble Ionic Compounds at 25°C**

Compound	K_{sp}	Compound	K_{sp}
Aluminum hydroxide [$Al(OH)_3$]	1.8×10^{-33}	Lead(II) chromate ($PbCrO_4$)	2.0×10^{-14}
Barium carbonate ($BaCO_3$)	8.1×10^{-9}	Lead(II) fluoride (PbF_2)	4.1×10^{-8}
Barium fluoride (BaF_2)	1.7×10^{-6}	Lead(II) iodide (PbI_2)	1.4×10^{-8}
Barium sulfate ($BaSO_4$)	1.1×10^{-10}	Lead(II) sulfide (PbS)	3.4×10^{-28}
Bismuth sulfide (Bi_2S_3)	1.6×10^{-72}	Magnesium carbonate ($MgCO_3$)	4.0×10^{-5}
Cadmium sulfide (CdS)	8.0×10^{-28}	Magnesium hydroxide [$Mg(OH)_2$]	1.2×10^{-11}
Calcium carbonate ($CaCO_3$)	8.7×10^{-9}	Manganese(II) sulfide (MnS)	3.0×10^{-14}
Calcium fluoride (CaF_2)	4.0×10^{-11}	Mercury(I) chloride (Hg_2Cl_2)	3.5×10^{-18}
Calcium hydroxide [$Ca(OH)_2$]	8.0×10^{-6}	Mercury(II) sulfide (HgS)	4.0×10^{-54}
Calcium phosphate [$Ca_3(PO_4)_2$]	1.2×10^{-26}	Nickel(II) sulfide (NiS)	1.4×10^{-24}
Chromium(III) hydroxide [$Cr(OH)_3$]	3.0×10^{-29}	Silver bromide ($AgBr$)	7.7×10^{-13}
Cobalt(II) sulfide (CoS)	4.0×10^{-21}	Silver carbonate (Ag_2CO_3)	8.1×10^{-12}
Copper(I) bromide ($CuBr$)	4.2×10^{-8}	Silver chloride ($AgCl$)	1.6×10^{-10}
Copper(I) iodide (CuI)	5.1×10^{-12}	Silver iodide (AgI)	8.3×10^{-17}
Copper(II) hydroxide [$Cu(OH)_2$]	2.2×10^{-20}	Silver sulfate (Ag_2SO_4)	1.4×10^{-5}
Copper(II) sulfide (CuS)	6.0×10^{-37}	Silver sulfide (Ag_2S)	6.0×10^{-51}
Iron(II) hydroxide [$Fe(OH)_2$]	1.6×10^{-14}	Strontium carbonate ($SrCO_3$)	1.6×10^{-9}
Iron(III) hydroxide [$Fe(OH)_3$]	1.1×10^{-36}	Strontium sulfate ($SrSO_4$)	3.8×10^{-7}
Iron(II) sulfide (FeS)	6.0×10^{-19}	Tin(II) sulfide (SnS)	1.0×10^{-26}
Lead(II) carbonate ($PbCO_3$)	3.3×10^{-14}	Zinc hydroxide [$Zn(OH)_2$]	1.8×10^{-14}
Lead(II) chloride ($PbCl_2$)	2.4×10^{-4}	Zinc sulfide (ZnS)	3.0×10^{-23}

Solubility Equilibria

- There are two other ways to express a substance’s solubility:
- ***Molar solubility (s)*** (mol/L) is the number of moles of solute dissolved in 1 L of a saturated solution.
- ***Solubility*** (g/L) is the number of grams of solute dissolved in 1 L of a saturated solution.
- Molar solubility = solubility / molar mass
- Solubility = molar solubility x molar mass

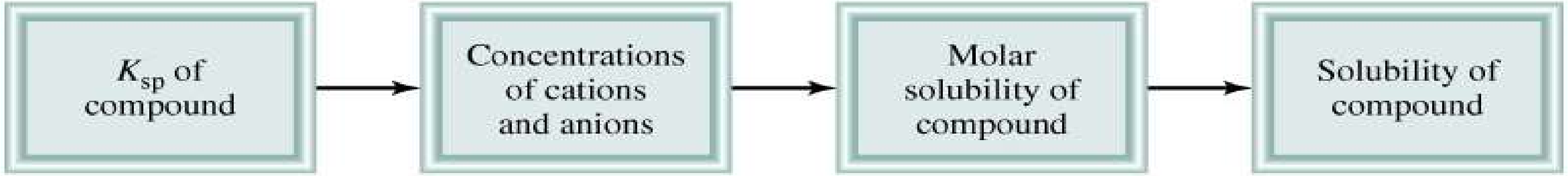


Solubility of compound

Molar solubility of compound

Concentrations of cations and anions

K_{sp} of compound



K_{sp} of compound

Concentrations of cations and anions

Molar solubility of compound

Solubility of compound

Solubility Equilibria

Example:

The solubility of calcium sulfate ($CaSO_4$) is found to be 0.67 g/L. Calculate the value of K_{sp} for calcium sulfate.?

$$CaSO_4(s) \rightleftharpoons Ca^{+2}\ (aq) + SO_4^{-2}\ (aq)$$

$[Ca^{+2}] = s$ $\quad$ $[SO_4^{-2}] = s$

$K_{sp} = [Ca^{+2}][SO_4^{-2}]$

$K_{sp} = s^2$

First, we calculate the number of moles of $CaSO_4$dissolved in 1 L of solution:

Molar solubility(s) = solubility / molar mass

$s = 0.67 /136.2 = 4.9x\ 10^{-3}$mol/L

$[Ca^{+2}]=4.9x10^{-3}$ and $[SO_4^{-2}]= 4.9x\ 10^{-3}$

$K_{sp} = [Ca^{+2}][SO_4^{-2}]$

$K_{sp} = (4.9x10^{-3})(4.9x10^{-3})$

$K_{sp} = 2.4x10^{-5}$

Solubility Equilibria

Example:

What is the solubility of $Cu(OH)_2$ in g/L if K_{sp} =2.2 x 10^{-20}

$$Cu(OH)(s) \rightleftharpoons Cu^{+2}\ (aq) + 2\ OH^{-}\ (aq)$$

$[Cu^{+2}]$= s $[OH^{-}]^{2}$= 2s

$K_{sp} = [Cu^{+2}][OH^{-}]^{2}$

$K_{sp} = (s)(2s)^{2}$

$K_{sp} = 4s^{3}$

$4s^{3}$= 2.2 x 10^{-20}

s^{3}= 5.5 x 10^{-21}

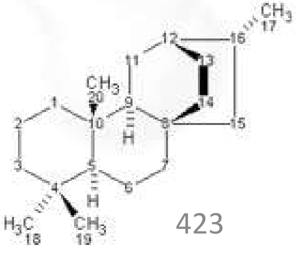
s= 1.8 x 10^{-7}M

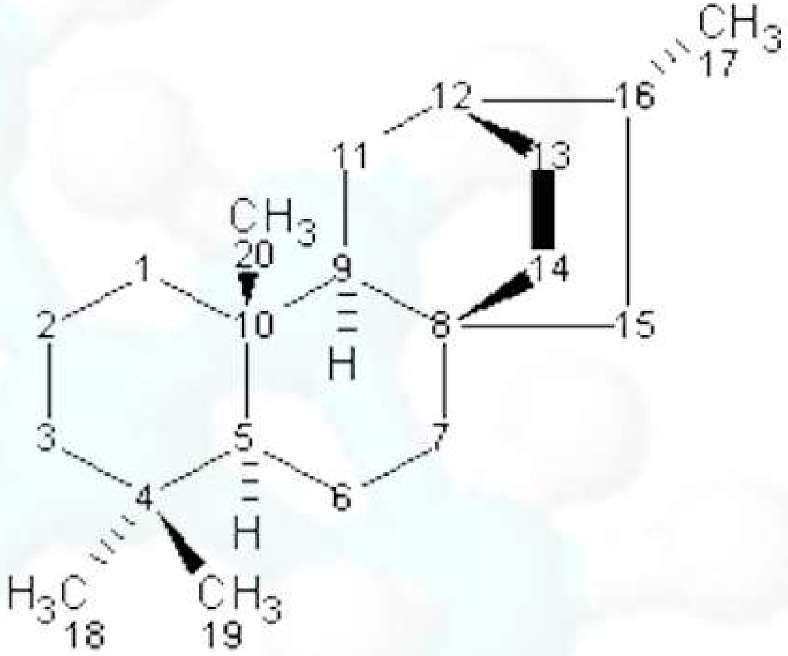
Solubility = molar solubility (s) x molar mass

solubility= 1.8 x 10^{-7}x 97.57

solubility= 1.8 x 10^{-5} g/L

Thank you





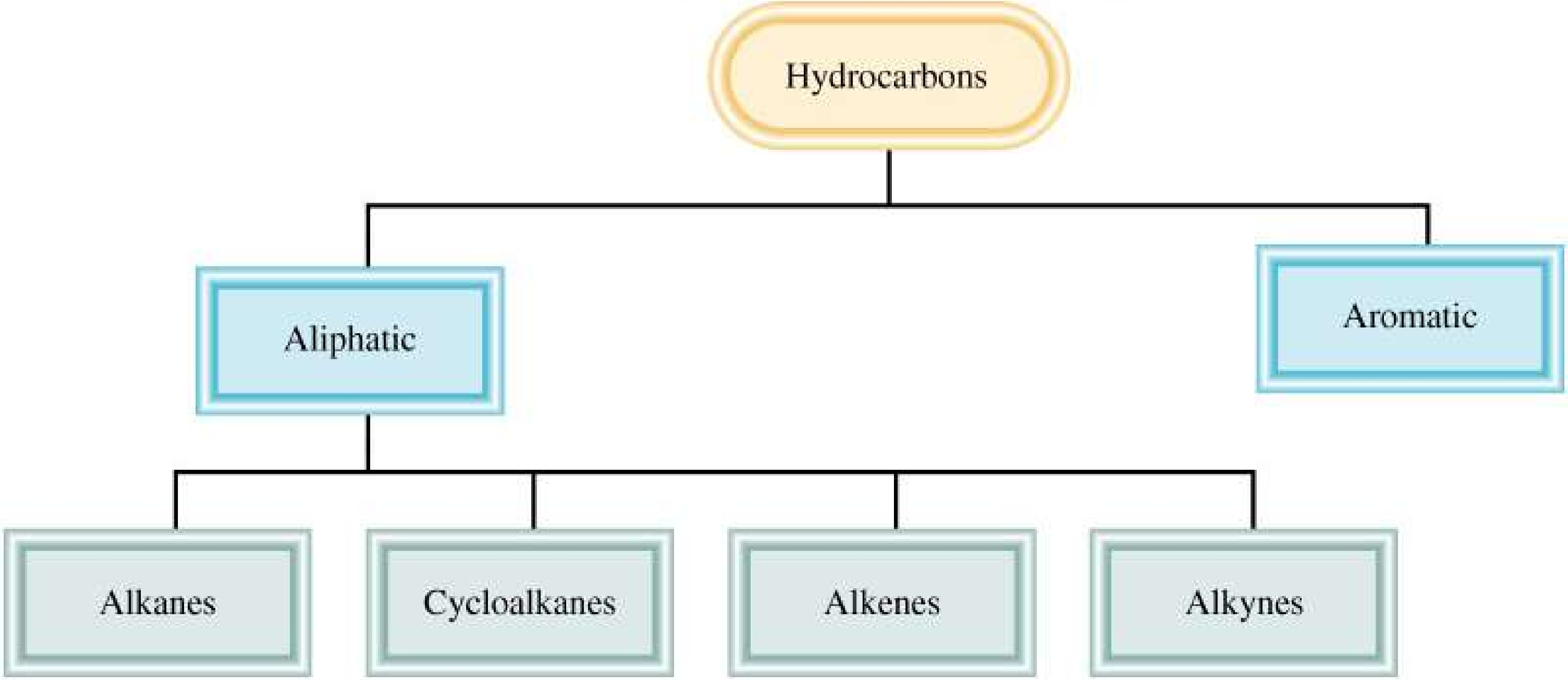
Chapter Twenty-four

Organic Chemistry

Classes of Organic Compounds

- The branch of chemistry that deals with carbon compounds is organic chemistry.
- Classes of organic compounds can be distinguished according to functional groups they contain.
- A functional group is a group of atoms that is largely responsible for the chemical behavior of the parent molecule.
- Most organic compounds are derived from a group of compounds known as hydrocarbons because they are made up of only hydrogen and carbon.
- Carbon has the ability to form long chains and ring structure.

Classes of Organic Compounds



Hydrocarbons

Aliphatic

Aromatic

Alkanes

Cycloalkanes

Alkenes

Alkynes

Classes of Organic Compounds

- Aliphatic hydrocarbons divided into:
  - **Alkanes:** Only single covalent bonds are present, general formula C_nH_{2n+2}, n = 1, 2, ….
  - **Cycloalkanes:** alkanes whose carbon atoms are joined in rings, general formula C_nH_{2n}, n = 3, 4, ….
  - **Alkenes:** contain at least one carbon-carbon double bond, general formula C_nH_{2n}, n = 2, 3 ….
  - **Alkynes:** contain at least one carbon-carbon triple bond, general formula C_nH_{2n-2}, n = 2, 3 ….

Classes of Organic Compounds

Example:

$C_{10}H_{22}$ is the formula of an:

[✓ a-alkane.] b-alkene. c-alkyne. d-aromatic hydrocarbon.

Aliphatic hydrocarbons

Alkanes

- ***Alkanes*** have the general formula C_nH_{2n+2} where *n* = 1,2,3,…
  - only single covalent bonds
  - ***saturated hydrocarbons*** because they contain the **maximum** number of hydrogen atoms that can bond with the number of carbon atoms in the molecule

CH_4	C_2H_6	C_3H_8
methane	ethane	propane

- Alkane Nomenclature:
- **IUPAC:** International Union of Pure and Applied Chemistry

Aliphatic hydrocarbons

Alkanes

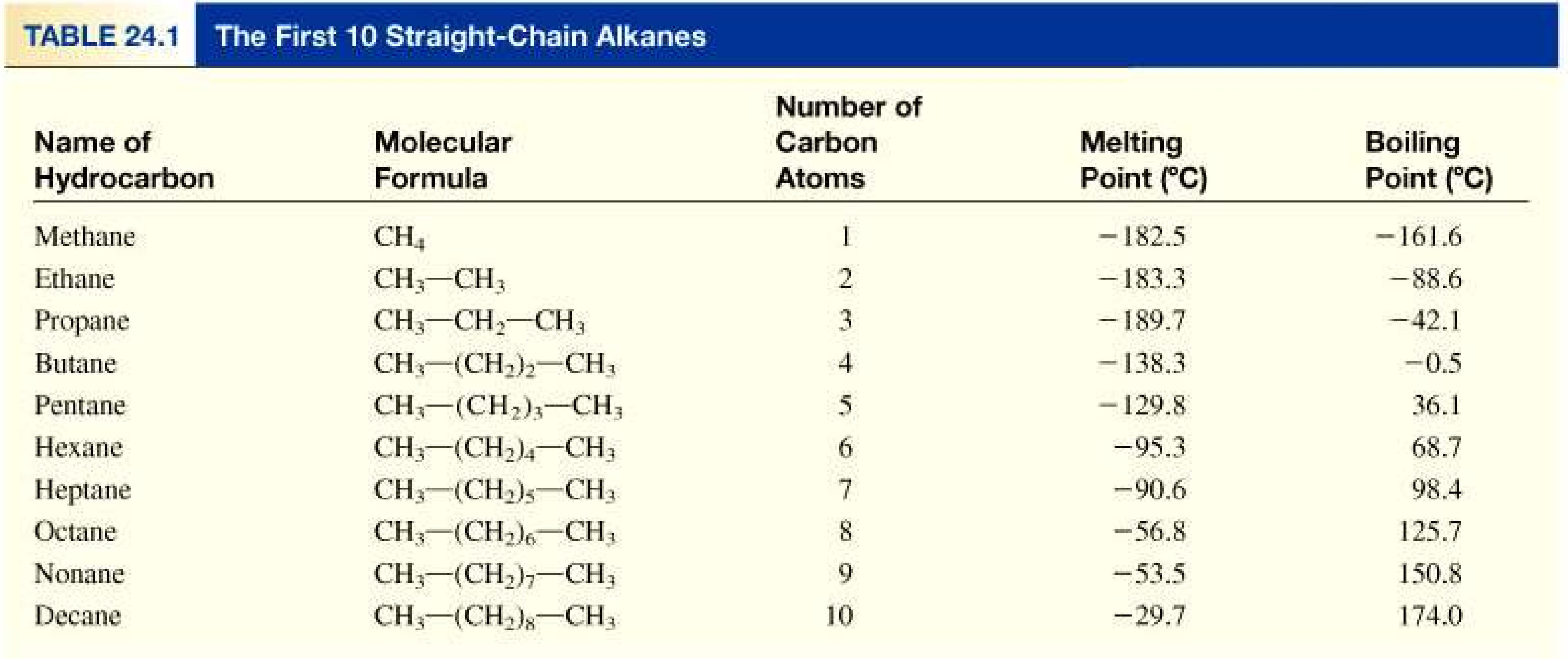


TABLE 24.1 The First 10 Straight-Chain Alkanes

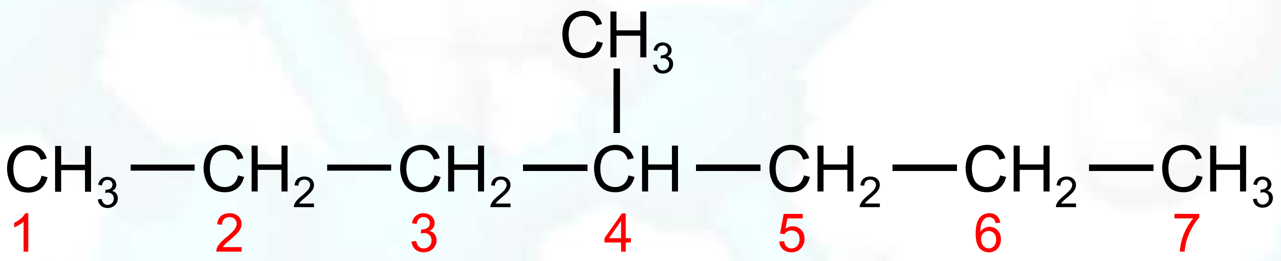
Name of Hydrocarbon	Molecular Formula	Number of Carbon Atoms	Melting Point (°C)	Boiling Point (°C)
Methane	CH_4	1	−182.5	−161.6
Ethane	$CH_3—CH_3$	2	−183.3	−88.6
Propane	$CH_3—CH_2—CH_3$	3	−189.7	−42.1
Butane	$CH_3—(CH_2)_2—CH_3$	4	−138.3	−0.5
Pentane	$CH_3—(CH_2)_3—CH_3$	5	−129.8	36.1
Hexane	$CH_3—(CH_2)_4—CH_3$	6	−95.3	68.7
Heptane	$CH_3—(CH_2)_5—CH_3$	7	−90.6	98.4
Octane	$CH_3—(CH_2)_6—CH_3$	8	−56.8	125.7
Nonane	$CH_3—(CH_2)_7—CH_3$	9	−53.5	150.8
Decane	$CH_3—(CH_2)_8—CH_3$	10	−29.7	174.0

Aliphatic hydrocarbons

Alkanes

1. The parent name of the hydrocarbon is that given to the longest continuous chain of carbon atoms in the molecule.

اختر أطول سلسلة ممكنة في الألكان



$$\overset{\displaystyle CH_3}{\overset{|}{}}$$
$$CH_3-CH_2-CH_2-CH(CH_3)-CH_2-CH_2-CH_3$$

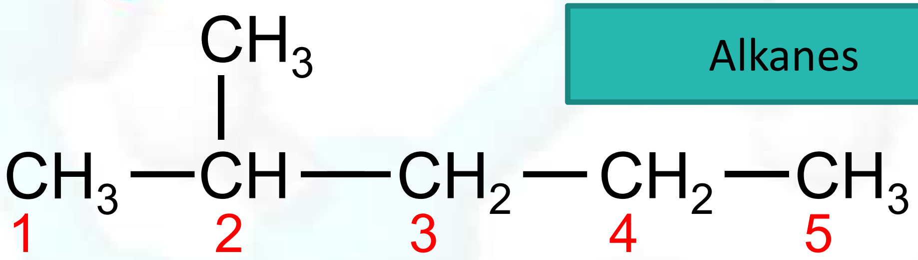
4-methylheptane

2. When one or more hydrogen atoms are replaced by other groups, the name of the compound must indicate the locations of carbon atoms where replacements are made. Number in the direction that gives the smaller numbers for the locations of the branches.

ابدأ بترقيم ذرات الكربون في السلسلة من الجهة الأقرب للتفرع.

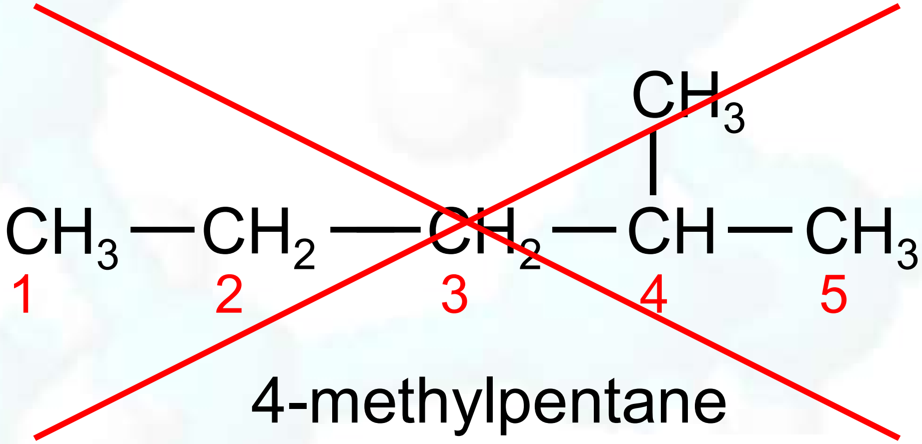
Aliphatic hydrocarbons

Alkanes



$$\underset{1}{CH_3}-\underset{2}{\overset{\overset{\displaystyle CH_3}{|}}{CH}}-\underset{3}{CH_2}-\underset{4}{CH_2}-\underset{5}{CH_3}$$

2-methylpentane



$$\xcancel{\underset{1}{CH_3}-\underset{2}{CH_2}-\underset{3}{CH_2}-\underset{4}{\overset{\overset{\displaystyle CH_3}{|}}{CH}}-\underset{5}{CH_3}}$$

4-methylpentane

Aliphatic hydrocarbons

Alkanes

3. An alkane less one hydrogen atom is an alkyl group.

CH_4 methane

CH_3 methyl

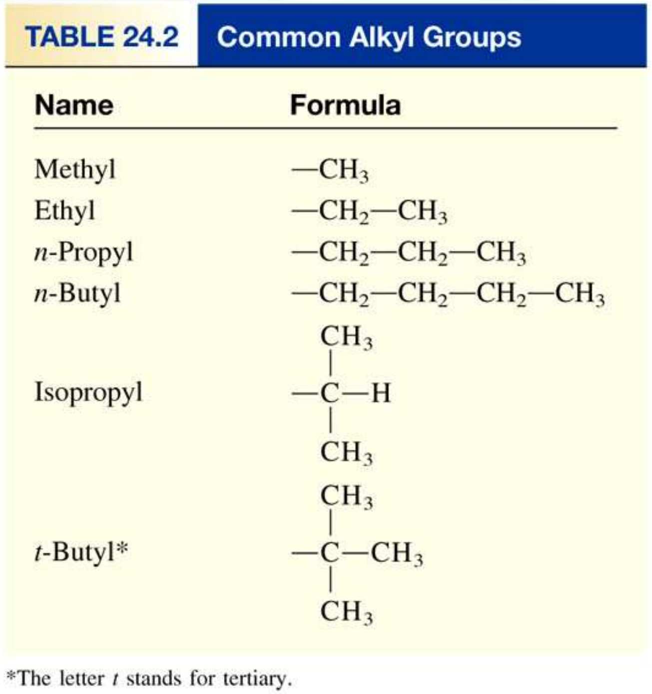


TABLE 24.2 **Common Alkyl Groups**

Name	Formula
Methyl	$-CH_3$
Ethyl	$-CH_2-CH_3$
n-Propyl	$-CH_2-CH_2-CH_3$
n-Butyl	$-CH_2-CH_2-CH_2-CH_3$
Isopropyl	$-\overset{\overset{\displaystyle CH_3}{\vert}}{\underset{\underset{\displaystyle CH_3}{\vert}}{C}}-H$
t-Butyl*	$-\overset{\overset{\displaystyle CH_3}{\vert}}{\underset{\underset{\displaystyle CH_3}{\vert}}{C}}-CH_3$

*The letter *t* stands for tertiary.

Aliphatic hydrocarbons

Alkanes

4. When there is more than one alkyl branch of the same kind present, we use a prefix such as di-, tri-, or tetra - with the name of the alkyl group.

في حال وجود أكثر من مجموعة متفرعة من نفس النوع، نستخدم prefixes **di-, tri-, tetra-**,

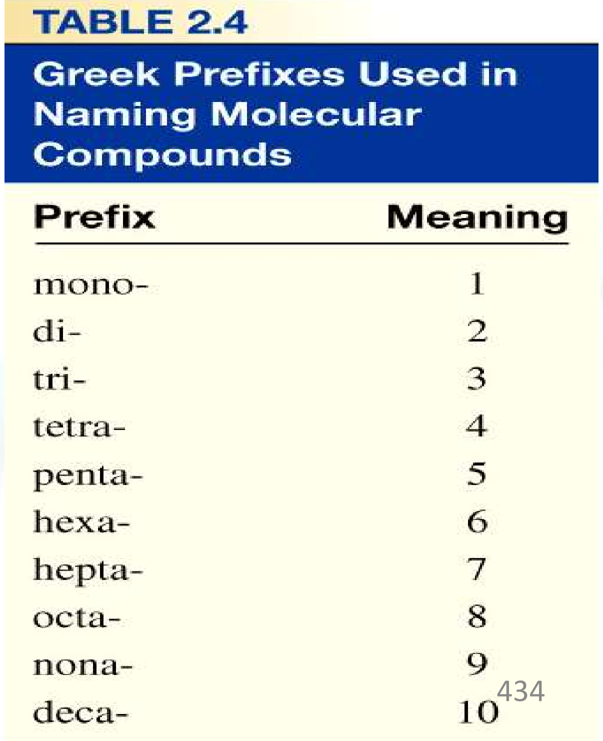


TABLE 2.4

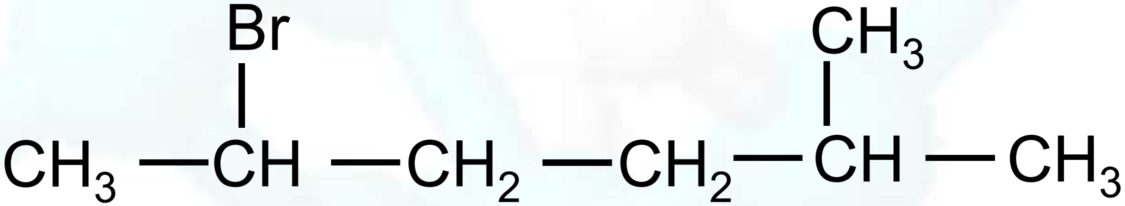
Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Aliphatic hydrocarbons

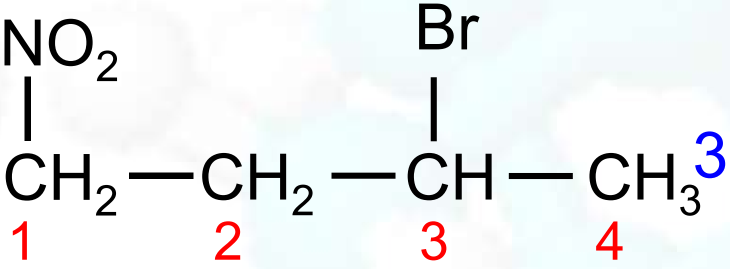
Alkanes

5. The substituent groups are listed alphabetically in the name, and the chain is numbered in the direction that gives the lowest number to the first substituted carbon atom.



$$CH_3-\overset{Br}{CH}-CH_2-CH_2-\overset{CH_3}{CH}-CH_3$$

2-bromo-5-methylhexane



$$\overset{NO_2}{CH_2}-CH_2-\overset{Br}{CH}-CH_3$$

3-bromo-1-nitrobutane

$$CH_3-\overset{Br}{CH}-\overset{NO_2}{CH}-CH_3$$

2-bromo-3-nitrobutane

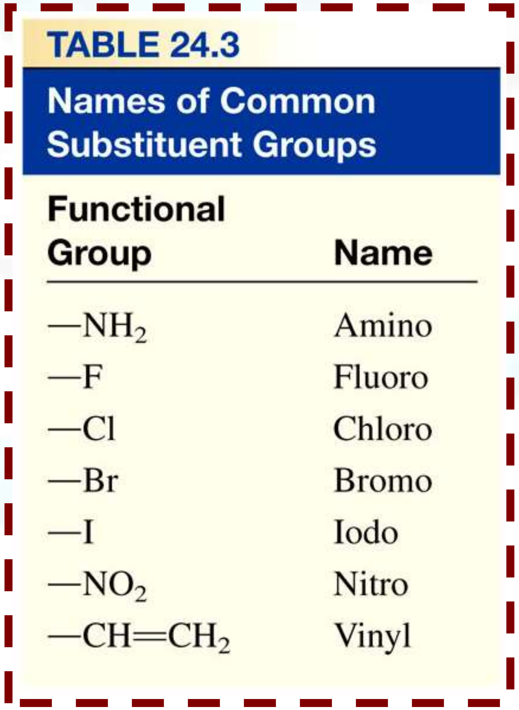


TABLE 24.3

Names of Common Substituent Groups

Functional Group	Name
$-NH_2$	Amino
$-F$	Fluoro
$-Cl$	Chloro
$-Br$	Bromo
$-I$	Iodo
$-NO_2$	Nitro
$-CH{=}CH_2$	Vinyl

Aliphatic hydrocarbons

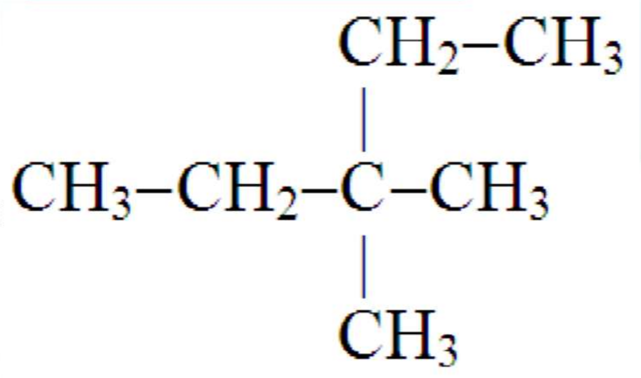
Alkanes

Example:

The systematic name for is

a- 1-ethyl-2-methylbutane

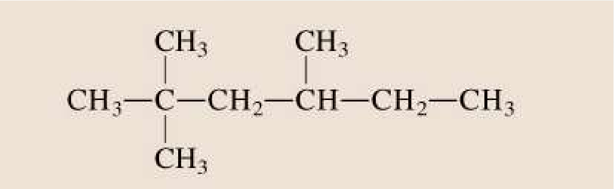
[✓ b- 3,3-dimethylpentane]



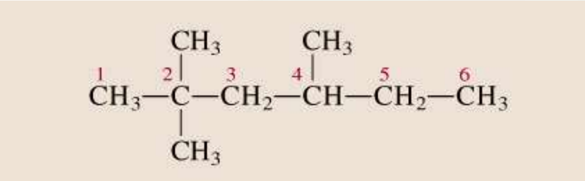
$$CH_3-CH_2-\underset{\underset{CH_3}{|}}{\overset{\overset{CH_2-CH_3}{|}}{C}}-CH_3$$

Example

Give the IUPAC name of the following compound:



$$CH_3-\underset{\underset{CH_3}{|}}{\overset{\overset{CH_3}{|}}{C}}-CH_2-\overset{\overset{CH_3}{|}}{CH}-CH_2-CH_3$$



$$\overset{1}{CH_3}-\underset{\underset{CH_3}{|}}{\overset{\overset{CH_3}{|}}{\overset{2}{C}}}-\overset{3}{CH_2}-\overset{\overset{CH_3}{|}}{\overset{4}{CH}}-\overset{5}{CH_2}-\overset{6}{CH_3}$$

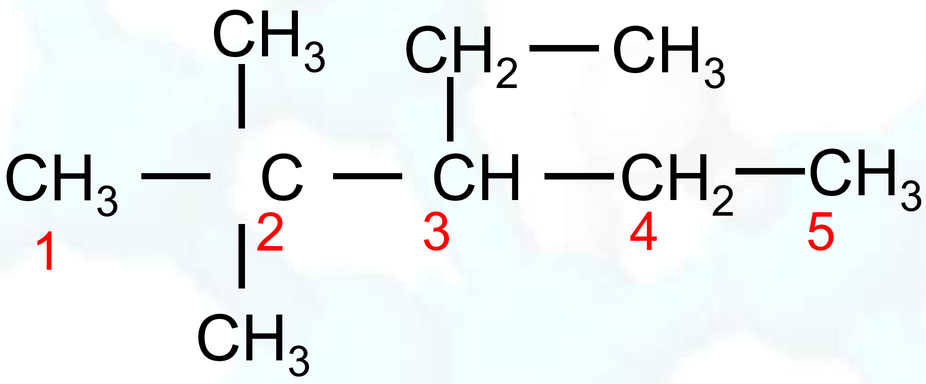
2,2,4-trimethylhexane

Aliphatic hydrocarbons

Alkanes

Example

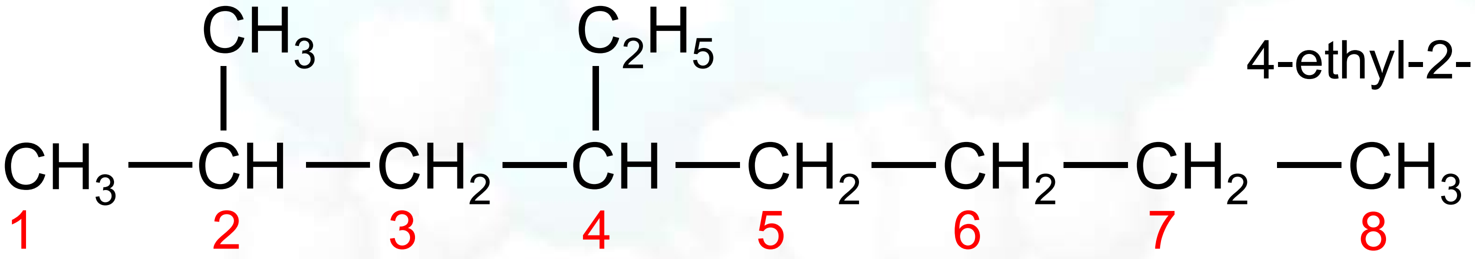
Write the structural formula of 3-ethyl-2,2-dimethylpentane



$$CH_3-\underset{\displaystyle CH_3}{\overset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}}-\overset{\displaystyle CH_2-CH_3}{\overset{|}{CH}}-CH_2-CH_3$$

Example

What is the IUPAC name of the following compound?



$$CH_3-\overset{\displaystyle CH_3}{\overset{|}{CH}}-CH_2-\overset{\displaystyle C_2H_5}{\overset{|}{CH}}-CH_2-CH_2-CH_2-CH_3$$

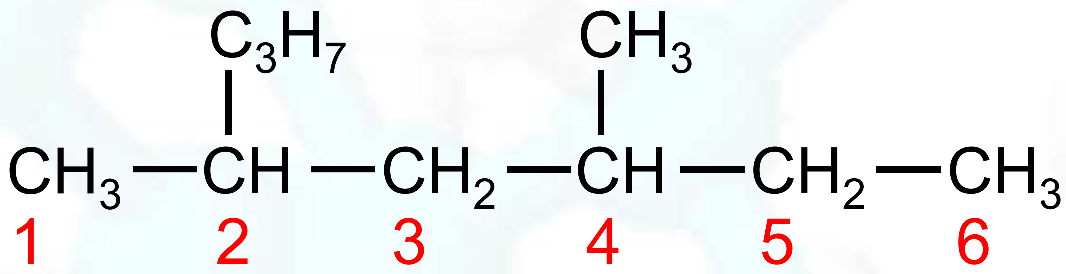
4-ethyl-2-methyloctane

Aliphatic hydrocarbons

Alkanes

Example

What is the structure of 4-methyl-2-propylhexane?



$$\begin{array}{ccccccccccc} & & C_3H_7 & & & & CH_3 & & & & \\ & & | & & & & | & & & & \\ CH_3 & - & CH & - & CH_2 & - & CH & - & CH_2 & - & CH_3 \end{array}$$

Aliphatic hydrocarbons

Alkanes

- Structural isomers: are molecules that have the same molecular formula but different structures.

# carbons	Name	# isomers
1	Methane	1
2	Ethane	1
3	Propane	1
4	Butane	2
5	Pentane	3
6	Hexane	5
7	Heptane	9

Aliphatic hydrocarbons

Alkanes

Example

How many structural isomers does pentane (C_5H_{12}) have?

a-1 b-2 [✓ c-3] d-4

Example

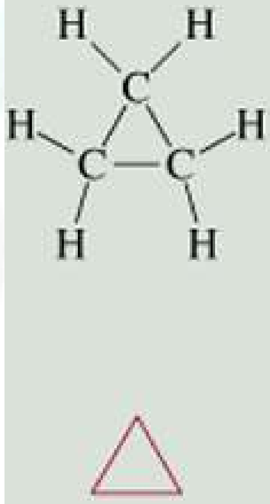
Which of the following hydrocarbons does **not** have isomers?

a-C_7H_{16} b-C_6H_{14} c-C_5H_{10} d-C_4H_8 [✓ e-C_3H_8]

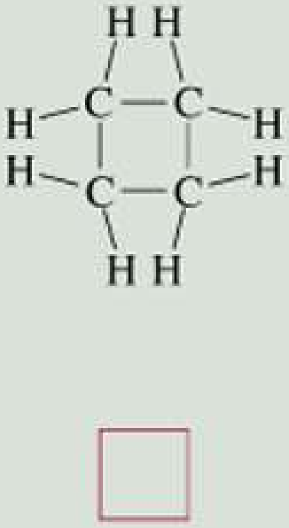
Aliphatic hydrocarbons

Cycloalkanes

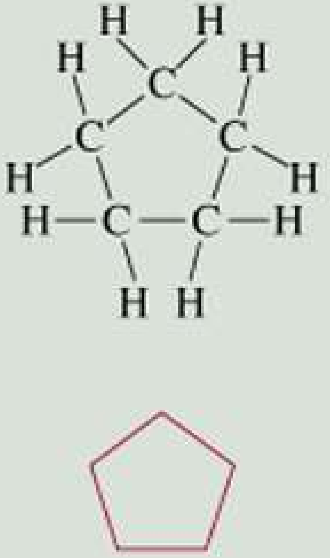
- Cycloalkanes are Alkanes whose carbon atoms are joined in rings.
- They have the general formula C_nH_{2n} where n = 3,4,…



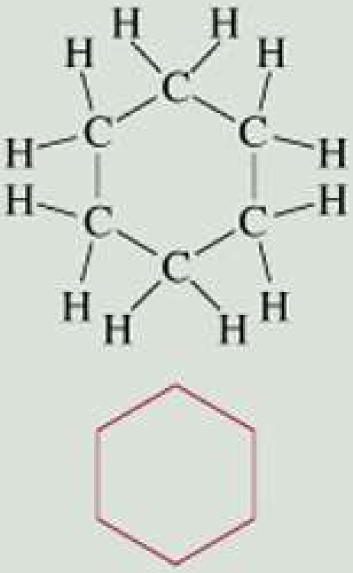
Cyclopropane



Cyclobutane



Cyclopentane



Cyclohexane

Aliphatic hydrocarbons

Alkenes

- Alkenes (also called olefins) are Alkanes contain at least one carbon-carbon double bond.
- They have the general formula C_nH_{2n} where n = 2,3,…
- The simplest alkene is C_2H_4 , ethylene $CH_2 = H_2C$

$$CH_2{=}CH{-}CH_2{-}CH_3$$

1-butene

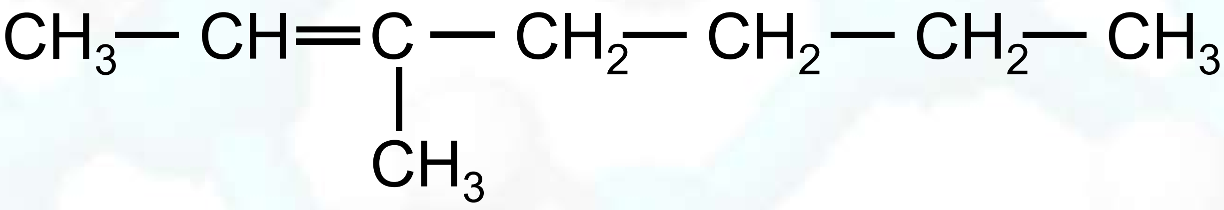
$$CH_3{-}CH{=}CH{-}CH_3$$

2-butene

- Alkene Nomenclature: same rules as alkane +
- The names of compounds containing C=C bonds end with –ene.
- The numbers in the names of alkenes refer to the lowest numbered carbon atom in the chain that is part of the C=C bond of the alkene.
- ابدأ بترقيم ذرات الكربون في السلسلة من الجهة الأقرب إلى الرابطة المزدوجة.

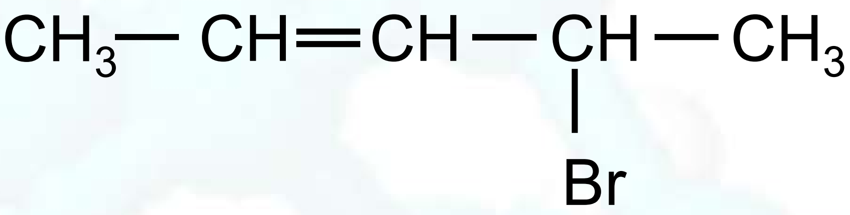
Aliphatic hydrocarbons

Alkenes



$$CH_3-CH=\underset{\displaystyle CH_3}{\underset{|}{C}}-CH_2-CH_2-CH_2-CH_3$$

3-methyl-2-heptene



$$CH_3-CH=CH-\underset{\displaystyle Br}{\underset{|}{CH}}-CH_3$$

4-bromo-2-pentene

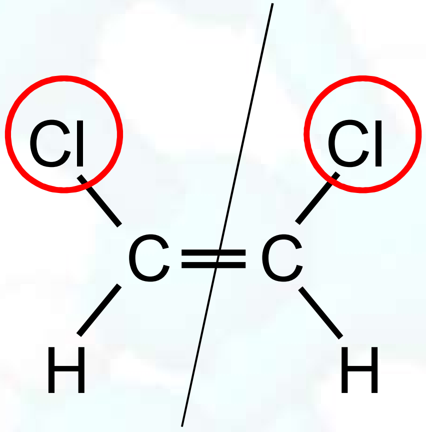
$CH_2=CH-CH_2-CH=CH_2$

1,4-pentadiene

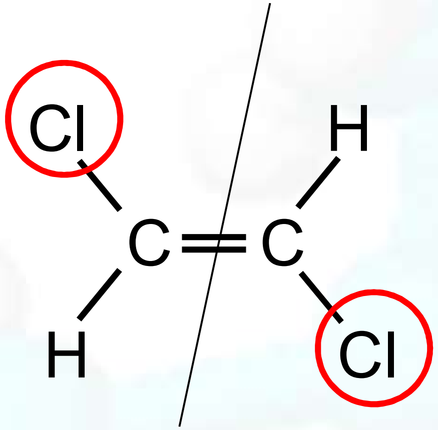
Aliphatic hydrocarbons

Alkenes

- Geometric Isomers of Alkenes: describing the orientation of functional group within a molecule.
- The terms *cis* on the same side, The terms *trans* on the other side.



cis-dichloroethylene



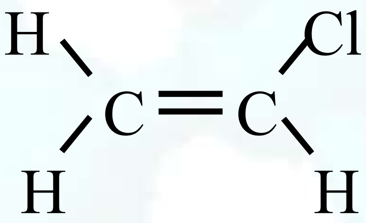
trans-dichloroethylene

Aliphatic hydrocarbons

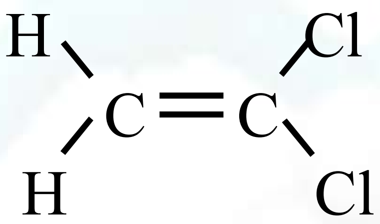
Alkenes

Example

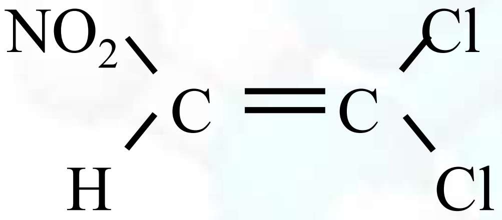
Which of the following compounds has geometrical isomer?



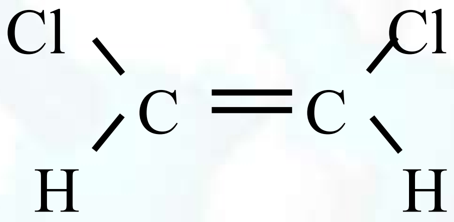
NO



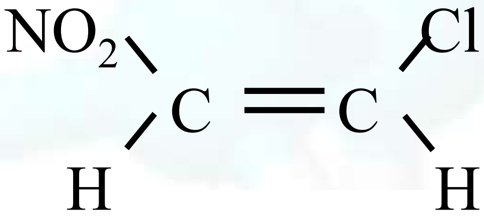
NO



NO



yes



yes

Aliphatic hydrocarbons

Alkenes

Example

For which of the compounds below are cis-trans isomers possible?

$CH_3CH{=}CH_2$ (1) $CH_3CH{=}CHCH_2CH_3$ (2) $CH_3CH{=}CHCH_3$ (3)

a-only 2   b-both 1 and 2   [✓ c-both 2 and 3]   d-all three

Example

Which of the following does <u>NOT</u> exhibit geometric isomerism?

a-4-octene   b-2-pentene   c-3-hexene   d-2-hexene   [✓ e-1-hexene]

Aliphatic hydrocarbons

Alkynes

- Alkynes are alkanes contain at least one carbon-carbon triple bond .
- They have the general formula C_nH_{2n-2} where n = 2,3,4,…
- Alkene Nomenclature: same rules as alkane +
- The names of compounds containing $C\equiv C$ bonds end with –yne.

$$CH \equiv C - CH_2 - CH_3$$

1-butyne

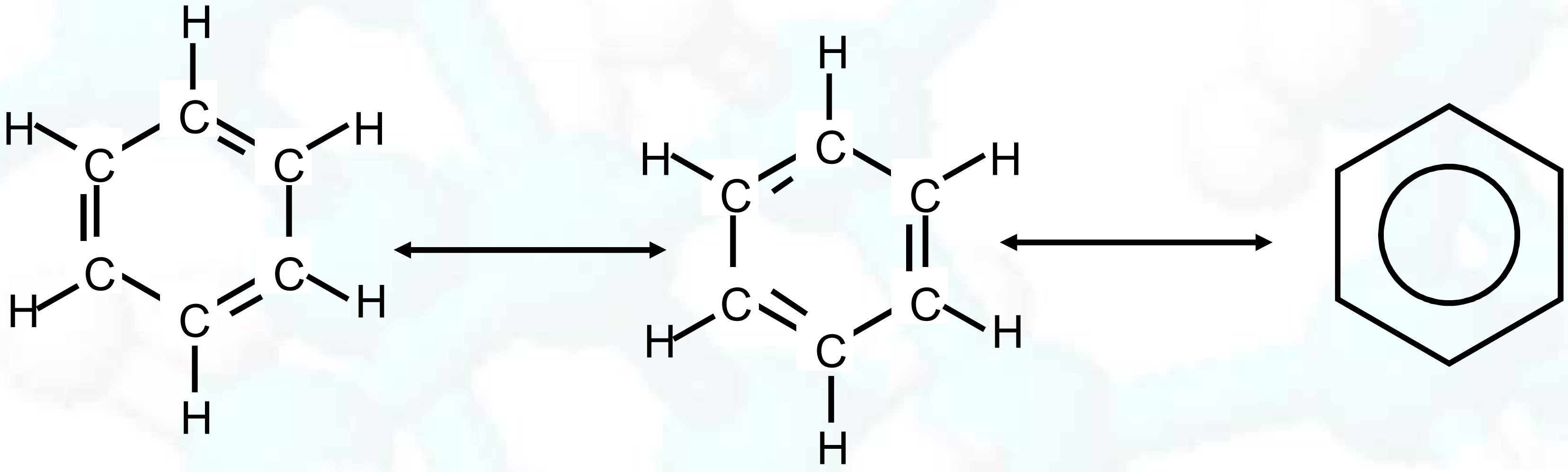
$$CH_3 - C \equiv C - CH_3$$

2-butyne

Classes of Organic Compounds

Aromatic

- Aromatic compound contain one or more benzene ring.



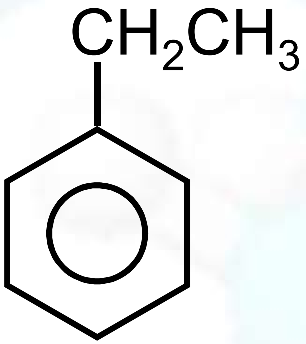
Classes of Organic Compounds

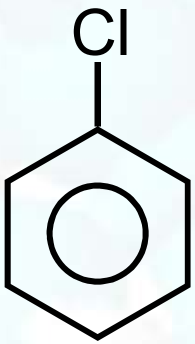
Aromatic

- Nomenclature: Same rules as before +

If one H atom has been replaced by another atom or a group of atom → the name of the atom or group then benzene

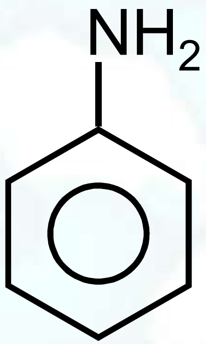
ethylbenzene

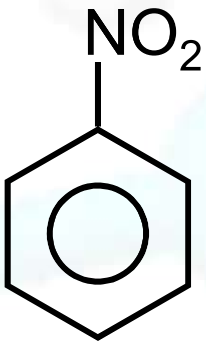




chlorobenzene

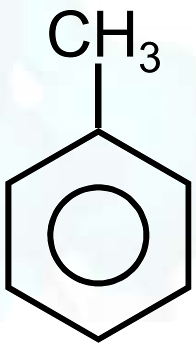
aminobenzene (aniline)





nitrobenzene

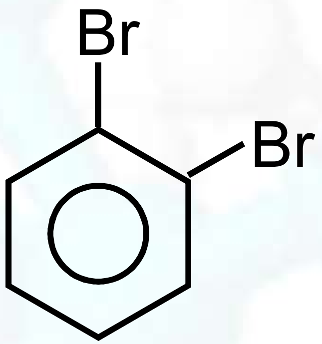
Toluene



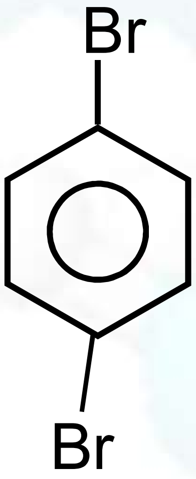
Classes of Organic Compounds

Aromatic

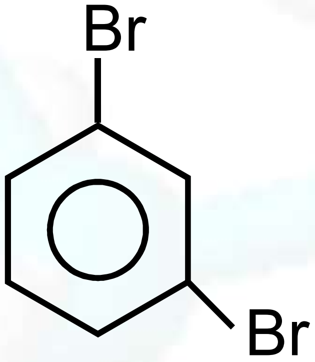
If more than one substituent is present, we must indicate the location of the second group relative to the first.



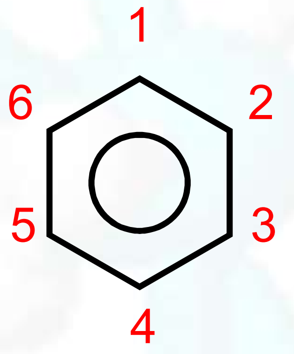
1,2-dibromobenzene
(o-dibromobenzene)



1,4-dibromobenzene
(p-dibromobenzene)



1,3-dibromobenzene
(m-dibromobenzene)



1,2 ortho o-

1,3 meta m-

1,4 para p-

Classes of Organic Compounds

Aromatic

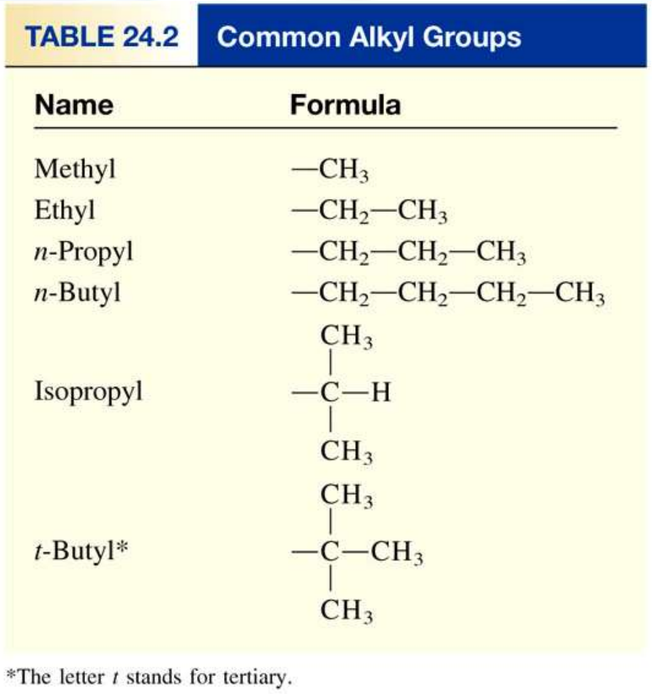


TABLE 24.2 Common Alkyl Groups

Name	Formula
Methyl	$-CH_3$
Ethyl	$-CH_2-CH_3$
n-Propyl	$-CH_2-CH_2-CH_3$
n-Butyl	$-CH_2-CH_2-CH_2-CH_3$
Isopropyl	$-\overset{CH_3}{\underset{CH_3}{C}}-H$
t-Butyl*	$-\overset{CH_3}{\underset{CH_3}{C}}-CH_3$

*The letter *t* stands for tertiary.

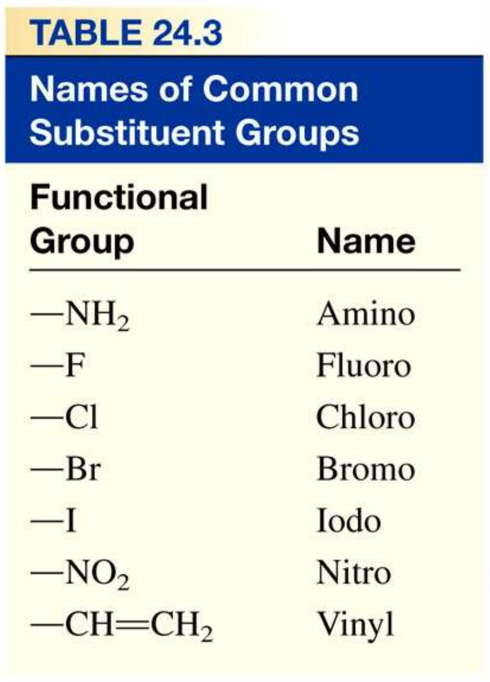


TABLE 24.3 Names of Common Substituent Groups

Functional Group	Name
$-NH_2$	Amino
$-F$	Fluoro
$-Cl$	Chloro
$-Br$	Bromo
$-I$	Iodo
$-NO_2$	Nitro
$-CH=CH_2$	Vinyl

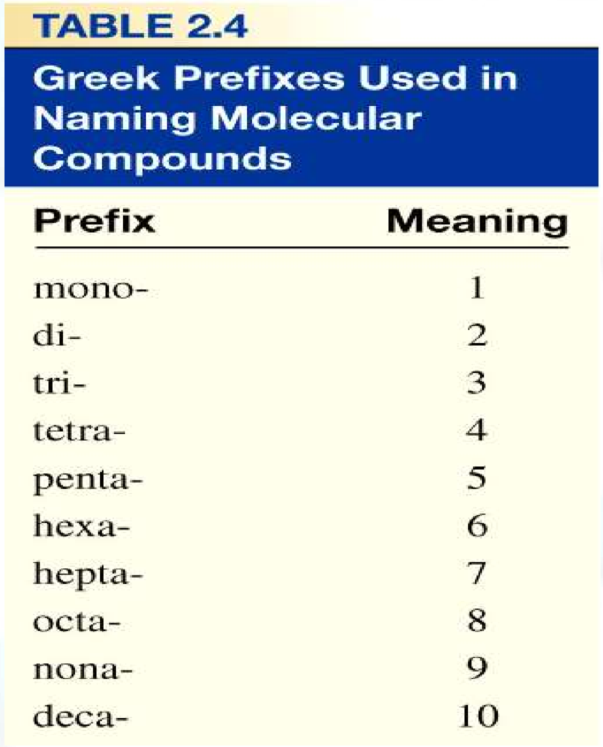
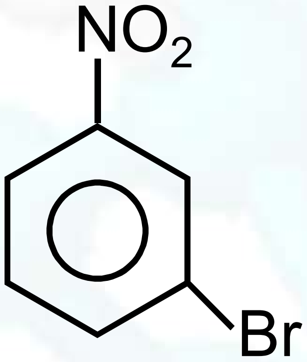


TABLE 2.4 Greek Prefixes Used in Naming Molecular Compounds

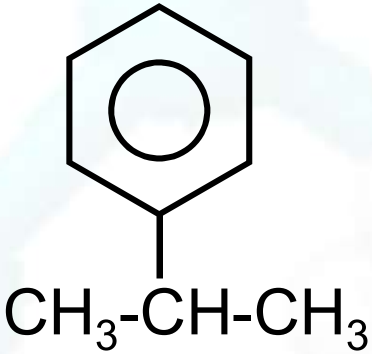
Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Classes of Organic Compounds

Aromatic



3-bromonitrobenzene
(*m*-bromonitrobenzene)



2-phenylpropane

The group containing benzene minus a hydrogen atom (C_6H_5) is called the phenyl group.

Classes of Organic Compounds

Aromatic

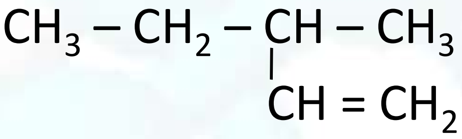
Example

Which one of these formulas is that of unsaturated hydrocarbon?

A. $CH_3 - CH_2 - CH_3$ [✓ B.$CH_3 - CH = CH_2$] C. $CH_3 - CH_2 - OH$

Example

Which of these is the systematic name for the compound represented below?



$$\begin{array}{l} CH_3 - CH_2 - \underset{\displaystyle CH = CH_2}{\underset{|}{CH}} - CH_3 \end{array}$$

A. 2-ethylbutane B. 3-methyl-1-hexene [✓ C. 3-methyl-1-pentene]

Hybridization

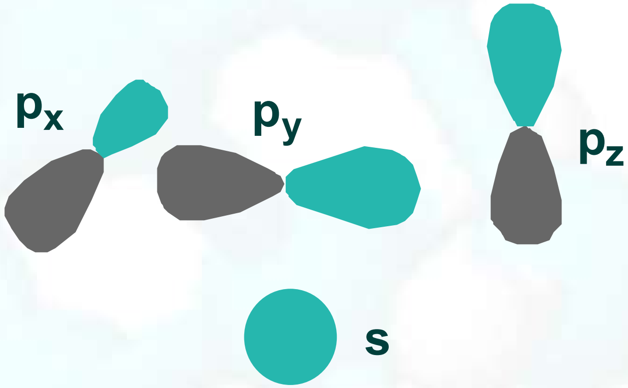
Hybridization: mixing of two or more atomic orbitals to form a new set of hybrid orbitals.

Hybridization is used to explain the formation of bonds.

In this course we will study three types of hybridization for carbon atom

SP^3 SP^2 SP

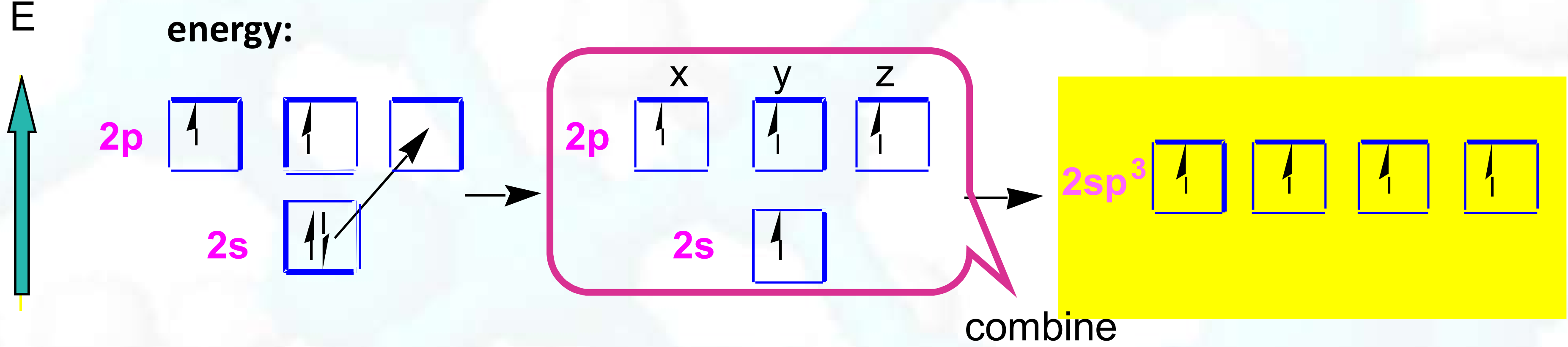
Remember



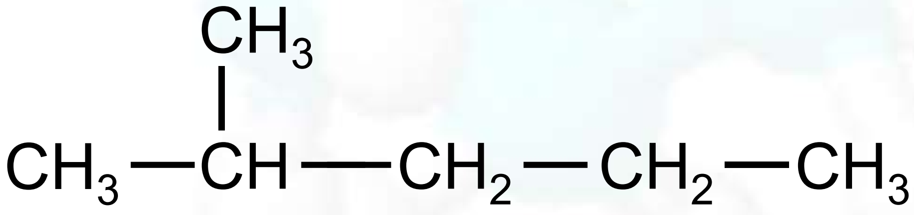
Hybridization

sp^3 hybridization

- **Promote one 2s electron into the vacant p-orbital.**
- **Combine (mix) all four orbitals to give four hybrid orbitals of equivalent energy:**



- **SP^3 hybridization always associated with saturated hydrocarbon compounds (i.e any carbon atom with 4 single bonds have SP^3 hybridization).**

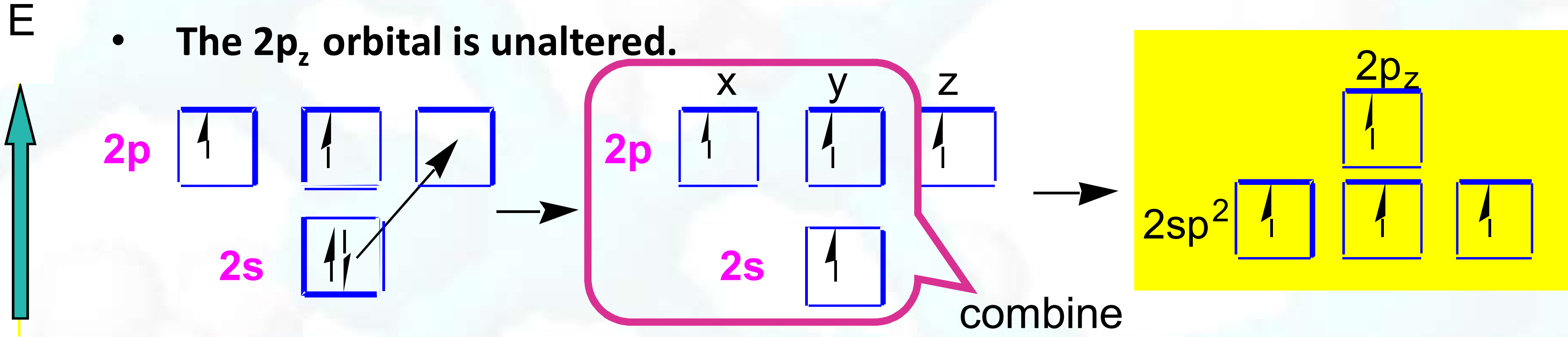


$$\overset{\displaystyle CH_3}{\overset{|}{}}\;CH_3-\overset{\overset{\displaystyle CH_3}{|}}{CH}-CH_2-CH_2-CH_3$$

$CH_3-CH_2-CH_2-CH_3$

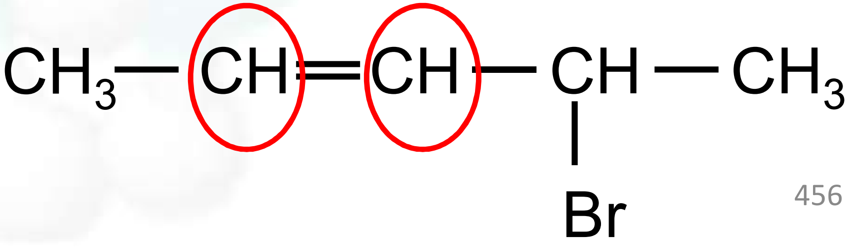
sp^2 hybridization

- **Promote one 2s electron into the vacant p-orbital.**
- **Combine (mix) the 2s, $2p_x$ and $2p_y$ orbitals to give three hybrid orbitals of equivalent energy**
- **The $2p_z$ orbital is unaltered.**



- **SP^2 hybridization always associated with unsaturated hydrocarbon compounds (i.e any carbon atom with 1 duple bonds have SP^2 hybridization).**

$CH_2{=}CH{-}CH_2{-}CH_3$

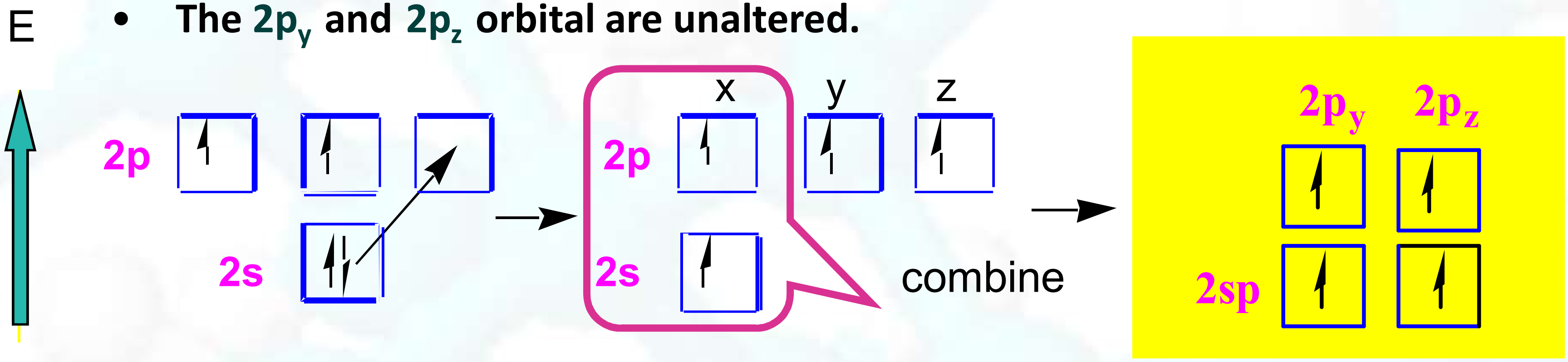


$CH_3{-}CH{=}CH{-}\underset{\displaystyle Br}{\underset{|}{CH}}{-}CH_3$

Hybridization

sp hybridization

- **Promote one 2s electron into the vacant p-orbital.**
- **Combine (mix) the 2s and $2p_x$ orbitals to give two hybrid orbitals of equivalent energy**
- **The $2p_y$ and $2p_z$ orbital are unaltered.**



- **SP hybridization always associated with unsaturated hydrocarbon compounds (i.e any carbon atom with 1 triple bonds or 2 duple bonds have SP hybridization).**

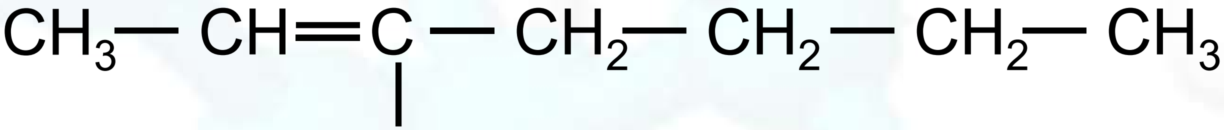
$CH_3-C\equiv C-CH_3$ $\qquad$ $CH_3-CH_2-CH=C=CH_2$

Hybridization

- **Sigma bond (σ) : the first bond made with any other atom.**

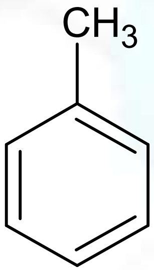
  Made from : hybridized orbitals
- **Pi bond (π): Any second or third bond made with any other atom**

  Made from : left over p orbital



$$CH_3-CH=C(-)-CH_2-CH_2-CH_2-CH_3$$

- **Sigma bond (σ) =23** **Pi bond (π)= 1**



Sigma bond (σ) =15 **Pi bond (π)= 3**

Classes of Organic Compounds

Functional Groups

Functional groups are responsible for most of the reactions of the parent compounds.

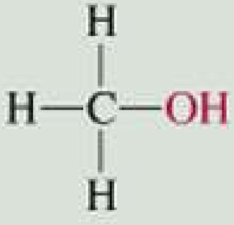
Functional groups are:

Alcohol , Ethers , Aldhyde , Ketones , Carboxylic acid , Esters , Amines , Aminoacid

Classes of Organic Compounds

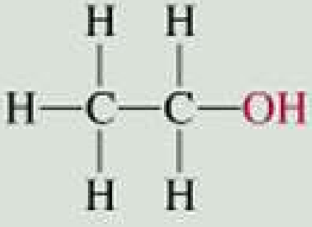
Functional Groups

Alcohol : contain the hydroxyl functional group, -OH.



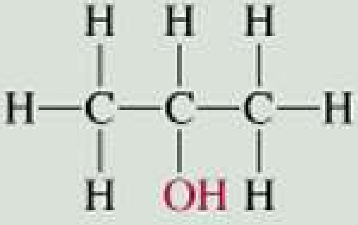
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & OH \\ & & | & & \\ & & H & & \end{array}$$

Methanol
(methyl alcohol)



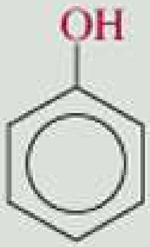
$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & OH \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

Ethanol
(ethyl alcohol)

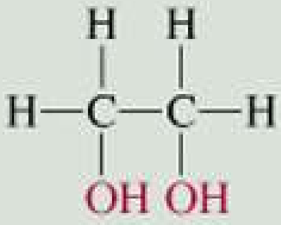


$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & \\ & & H & & OH & & H & & \end{array}$$

2-Propanol
(isopropyl alcohol)



Phenol



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & H \\ & & | & & | & & \\ & & OH & & OH & & \end{array}$$

Ethylene glycol

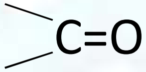
Classes of Organic Compounds

Functional Groups

Ethers: contain the R-O-R’ linkage, where R and R’ are a hydrocarbon (aliphatic or aromatic) group.

CH_3OCH_3 Dimethyl ether $C_2H_5OC_2H_5$ Diethyl ether

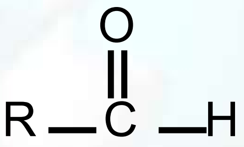
Aldehydes and Ketones: The functional group in these compounds is the carbonyl group.



$>C=O$

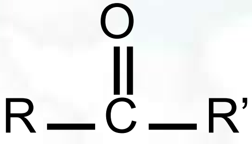
In an aldehyde at least one hydrogen atom is bonded to the carbon in the carbonyl group. In a ketone, the carbon atom in the carbonyl group is bonded to two hydrocarbon groups.

- **aldehydes** have the general formula



$R-\overset{\overset{\displaystyle O}{\|}}{C}-H$

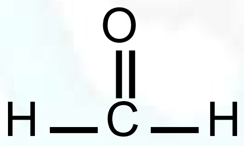
- **ketones** have the general formula



$R-\overset{\overset{\displaystyle O}{\|}}{C}-R'$

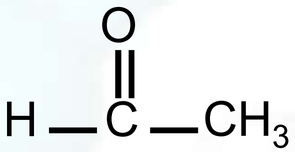
Classes of Organic Compounds

Functional Groups



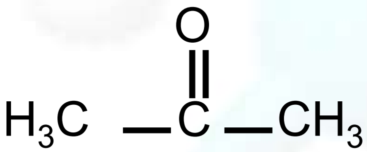
$$H-\overset{\overset{O}{\|}}{C}-H$$

formaldehyde



$$H-\overset{\overset{O}{\|}}{C}-CH_3$$

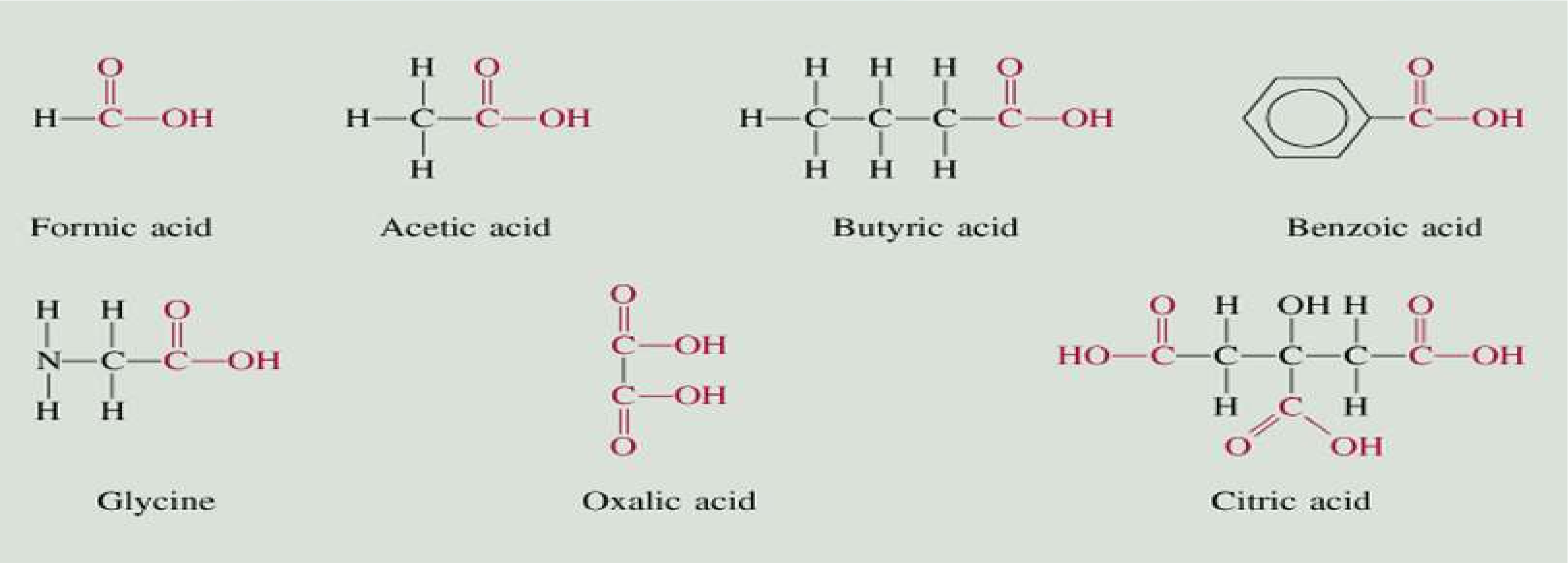
acetaldehyde



$$H_3C-\overset{\overset{O}{\|}}{C}-CH_3$$

acetone

Carboxylic Acids: acids that contain the carboxyl group, -COOH.



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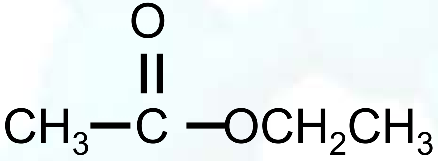
Esters: have the general formula R’COOR, where R’ can be H or a hydrocarbon group and R is a hydrocarbon group.



$$CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-OCH_2CH_3$$

ethyl acetate

$CH_3CH_2CH_2COOCH_3$

methyl butyrate

Amines: are organic bases having the general formula R_3N, where R may be H or a hydrocarbon group.

CH_3NH_2

Methyl amine

$CH_3CH_2NH_2$

Ethyl amine

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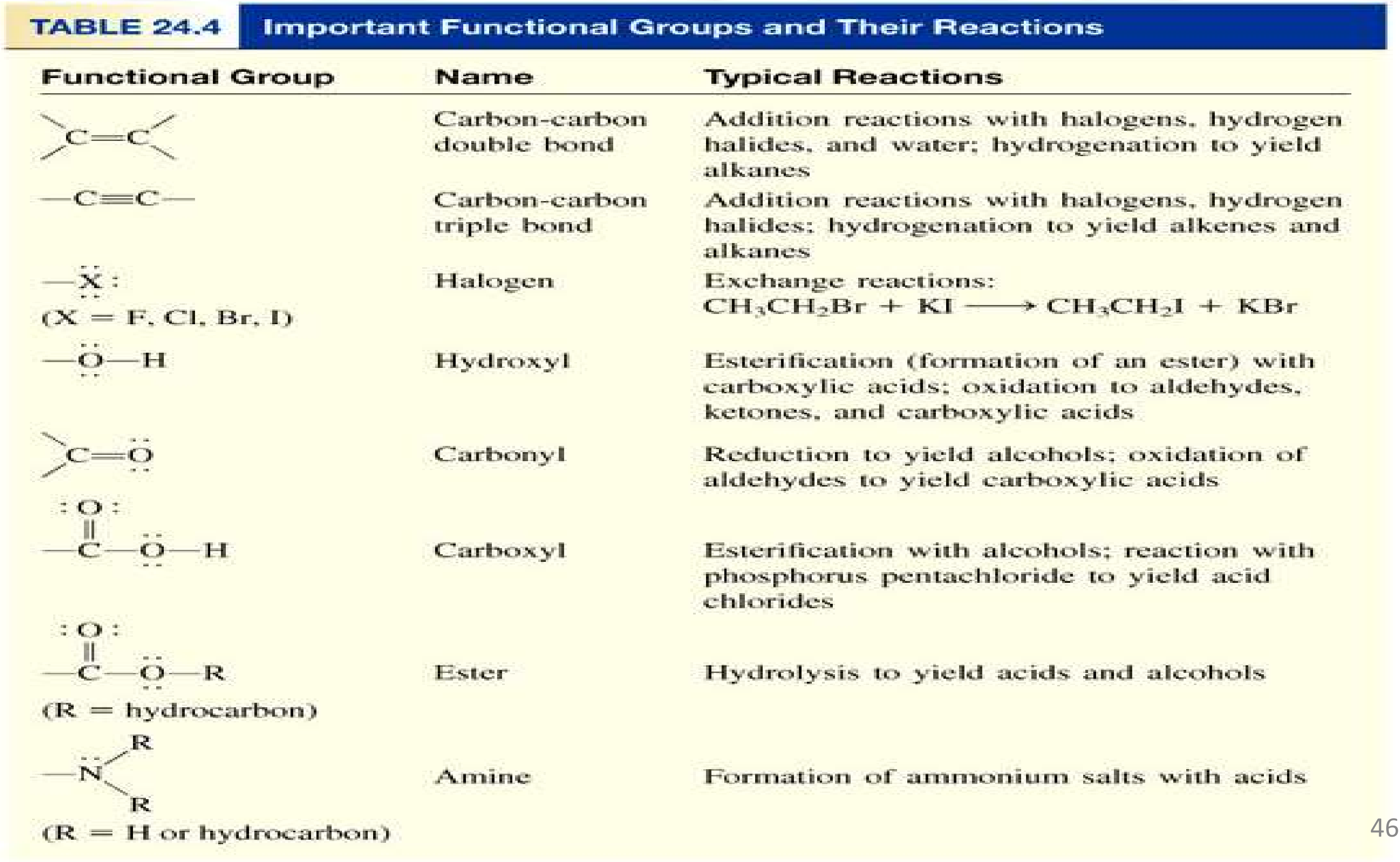

TABLE 24.4 **Important Functional Groups and Their Reactions**

Functional Group	Name	Typical Reactions
$>C=C<$	Carbon-carbon double bond	Addition reactions with halogens, hydrogen halides, and water; hydrogenation to yield alkanes
$-C\equiv C-$	Carbon-carbon triple bond	Addition reactions with halogens, hydrogen halides; hydrogenation to yield alkenes and alkanes
$-\underset{..}{\ddot{X}}:$ $(X = F, Cl, Br, I)$	Halogen	Exchange reactions: $CH_3CH_2Br + KI \longrightarrow CH_3CH_2I + KBr$
$-\underset{..}{\ddot{O}}-H$	Hydroxyl	Esterification (formation of an ester) with carboxylic acids; oxidation to aldehydes, ketones, and carboxylic acids
$>C=\underset{..}{\ddot{O}}$	Carbonyl	Reduction to yield alcohols; oxidation of aldehydes to yield carboxylic acids
$-\overset{\overset{:O:}{\Vert}}{C}-\underset{..}{\ddot{O}}-H$	Carboxyl	Esterification with alcohols; reaction with phosphorus pentachloride to yield acid chlorides
$-\overset{\overset{:O:}{\Vert}}{C}-\underset{..}{\ddot{O}}-R$ $(R = \text{hydrocarbon})$	Ester	Hydrolysis to yield acids and alcohols
$-\ddot{N}\genfrac{}{}{0pt}{}{\diagup R}{\diagdown R}$ $(R = \text{H or hydrocarbon})$	Amine	Formation of ammonium salts with acids

Thank you